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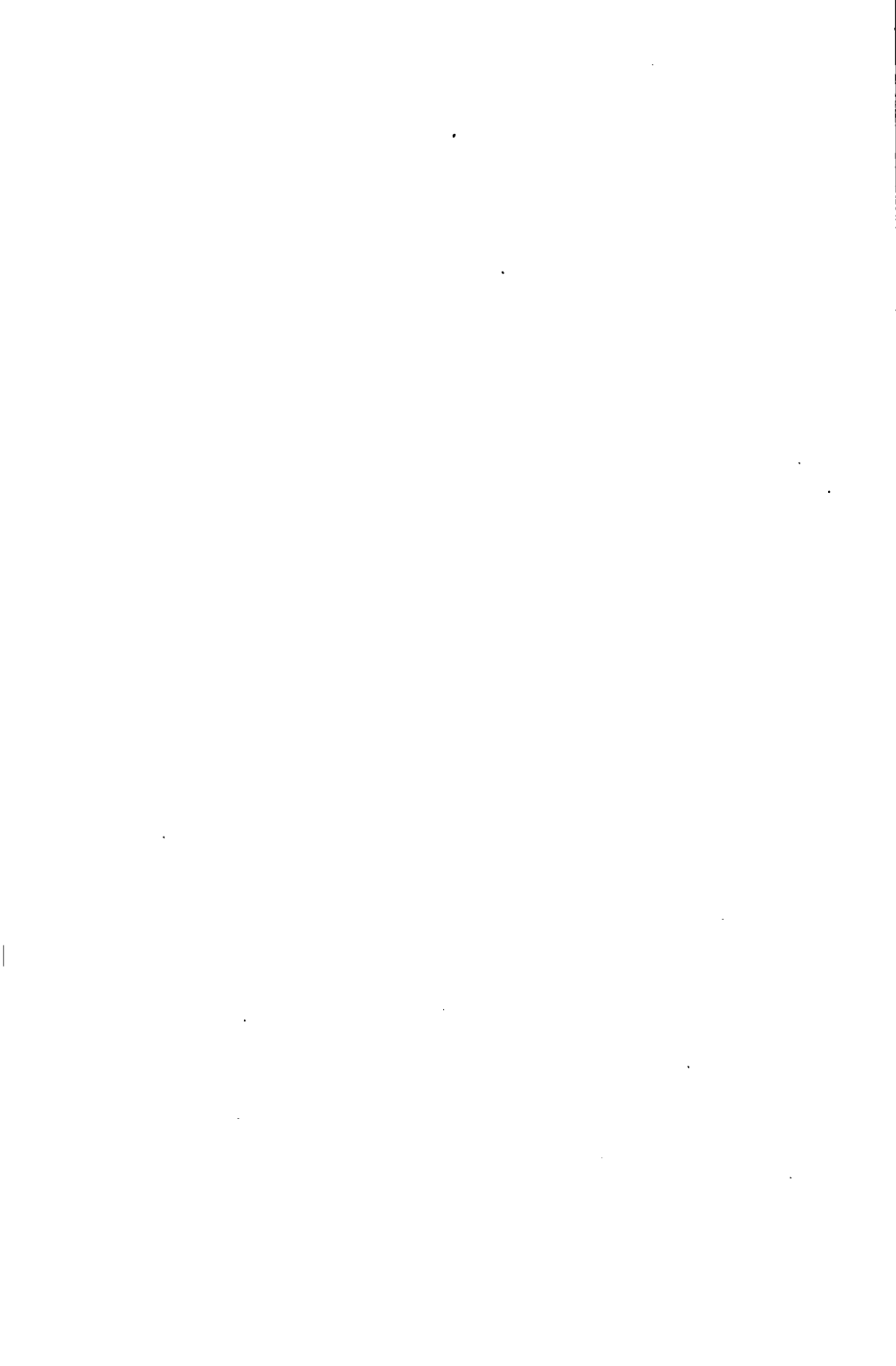
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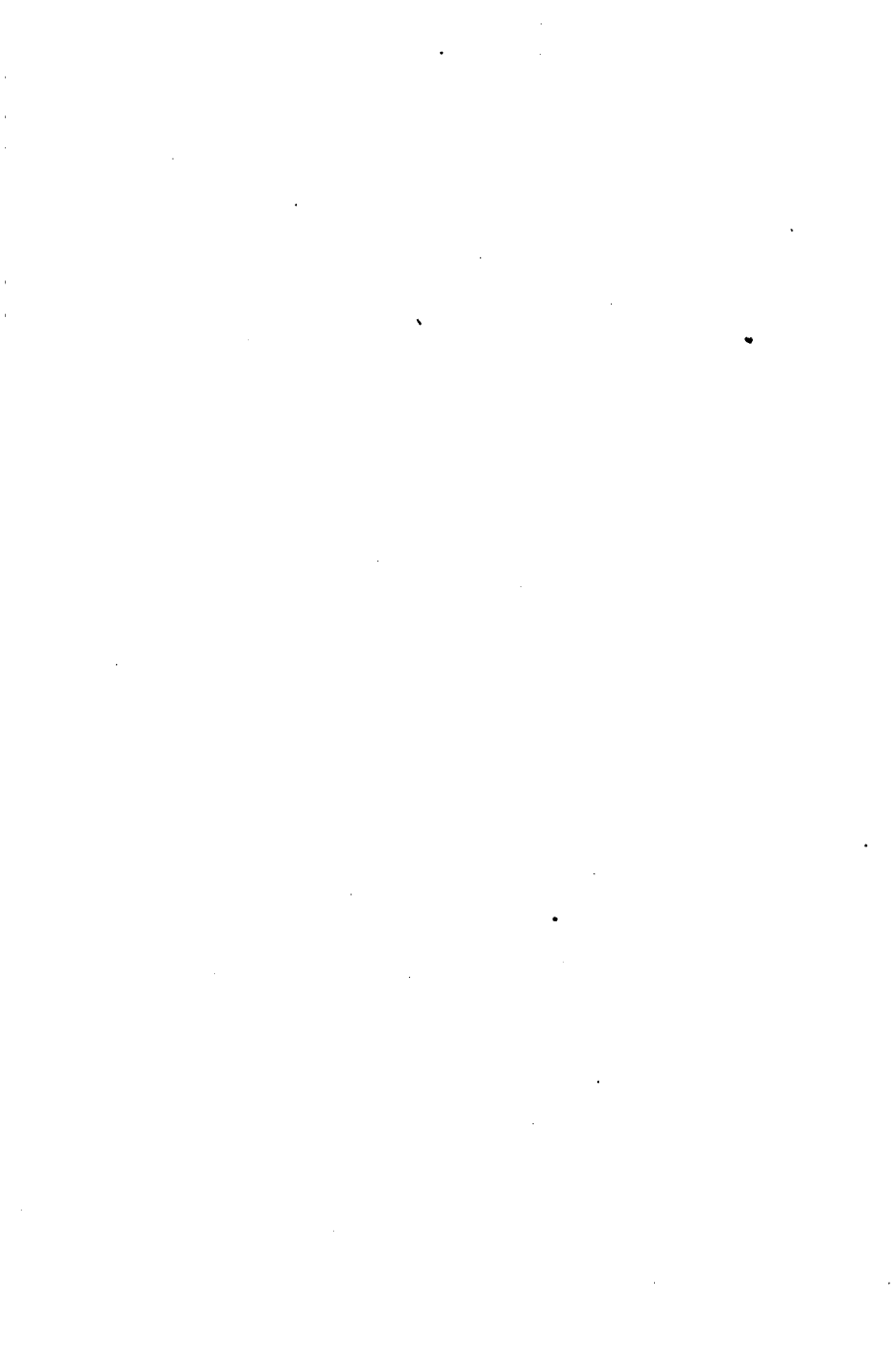
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LABORATORY WORK

IN

CHEMISTRY

*A SERIES OF EXPERIMENTS IN GENERAL  
INORGANIC CHEMISTRY*

BY

EDWARD H. KEISER

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LAB. WORK IN CHEM.

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## PREFACE.

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THE directions for laboratory experiments contained in this book have been arranged for the use of students that are following a course of lectures or recitations on general chemistry. These laboratory exercises are intended to illustrate and to be supplementary to the work of the class room. It is not intended that this book shall take the place of the personal instruction of the teacher; on the contrary, its object is to facilitate this work of the instructor where the number of students in the class is large.

The laboratory work and the lectures or recitations should go hand in hand. Before a subject is taken up in the laboratory it ought to be briefly outlined in the class room, and the object of the experiments that are to be made ought to have been clearly brought into the minds of the students. At the same time they must not be told everything that they are expected to observe; they must be given opportunity to acquire the art of observing accurately and of describing what they see. After all the experiments on a given subject have been made, a full discussion of the results is in place. Then the instructor should see to it that the students have observed everything that was to be seen, that proper conclusions have been drawn from the results of the experiments, and that the connections between these results and the general laws and principles of the science are clearly brought out.

Some of the experiments illustrating important facts and laws of chemistry are, on account of the expensive apparatus or chemicals required, or because of the skill and care necessary in their execution, of such a nature that they cannot well be done by each individual member of a large class; nevertheless they are of such importance that each student ought to have an accurate knowledge of them. Experiments of this kind have been called Laboratory Demonstrations in the text; and it is intended that these experiments shall be carried out in the presence of the whole class by one or two of the more skillful students, working under the immediate supervision of the instructor. After each such demonstration, opportunity ought to be given to the class to ask questions and to discuss the results.

In doing the work outlined in these pages, the student is expected to keep a full record of each experiment (laboratory demonstrations included) in a notebook and to conscientiously do the work indicated. The notebook ought to be examined and corrected from time to time by the instructor. Each student should own a good text-book of chemistry and ought to have ready access to the chemical dictionaries and works of reference.

In conclusion I desire to express my thanks to Dr. E. P. Kohler for valuable suggestions made to me in regard to the directions for some of the experiments, and also to Professor Ira Remsen for permission to use the directions for several experiments that are given in his text-books.

I am also under obligations to the American Book Co., publishers of Storer and Lindsay's Manual of Chemistry, for the privilege of using some of the illustrations contained in that text-book.

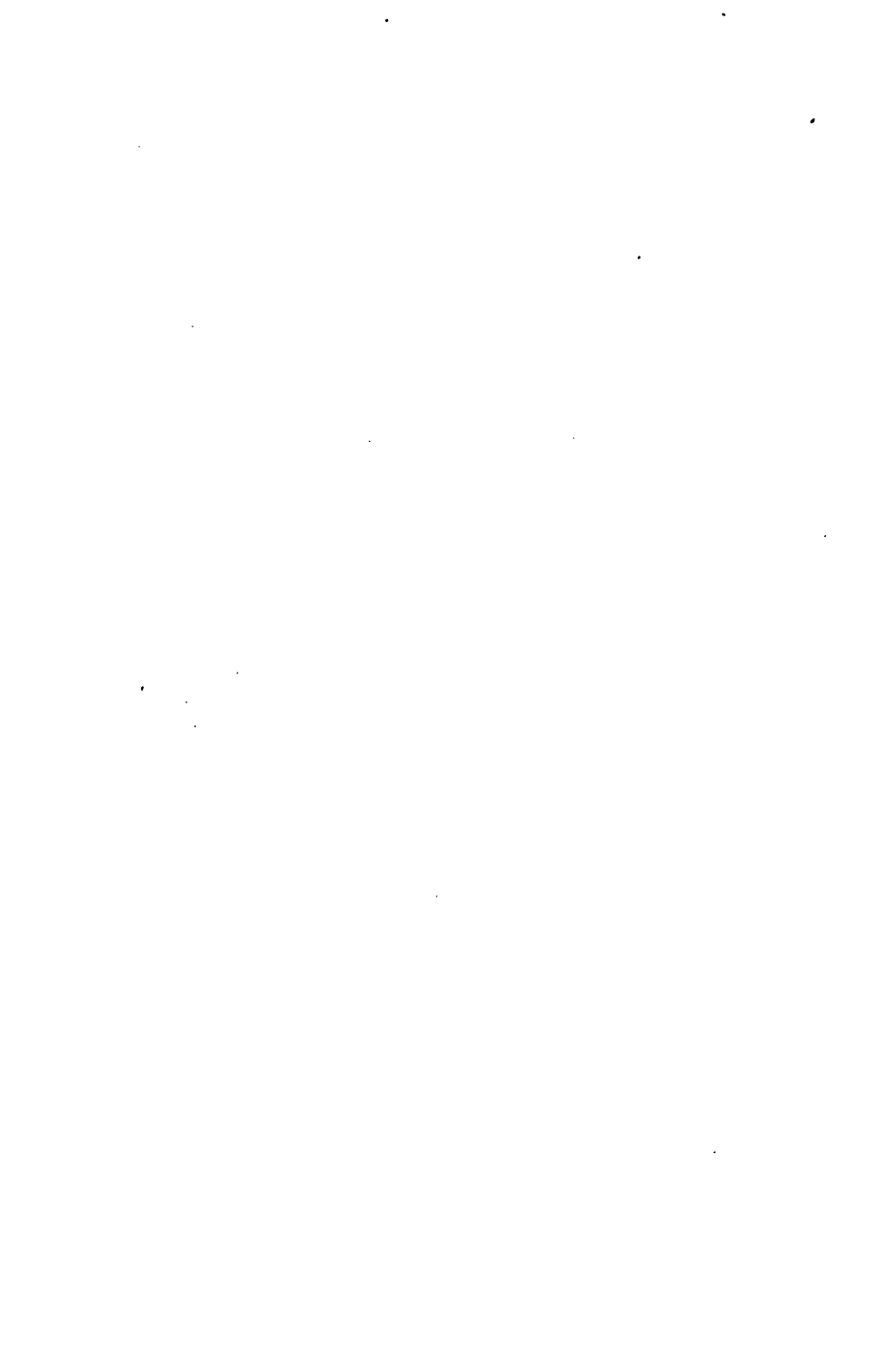
## SUGGESTIONS TO STUDENTS

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BEFORE beginning an experiment, read carefully all the directions for the work. Arrange the apparatus as directed, or if you desire to use a modified form of apparatus, consult the instructor before proceeding. In your notebook describe the apparatus and materials that you use and what you have done with them. A simple sketch of apparatus is very useful. Describe all the changes that occur, and answer the questions that are given in the text.

In generating gases make sure that all joints and stoppers of the apparatus are tight. If the generating vessels are to be heated, it is necessary to use thin-walled flasks. Thick glass bottles almost always crack when heated. Small quantities of liquids are heated in test tubes. Solids are heated in ignition tubes, and all glassware must be dry on the outside before heating. Always keep your apparatus and desk clean and neat.

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# TABLE OF CONTENTS.

	PAGE
Exercises in the Construction of Apparatus . . . . .	1
Changes in Matter, Chemical and Physical Changes . . . . .	3
Matter and Energy . . . . .	7
Mechanical Mixtures and Chemical Compounds . . . . .	8
Homogeneous Matter, Elements, Compounds, Physical Mixtures . . . . .	9
Chemical Balance . . . . .	10
Measurement of Liquids . . . . .	10
The Air . . . . .	11
Measurement of Gases . . . . .	14
Oxygen . . . . .	16
The Conservation of Matter . . . . .	20
Laws of Conservation of Matter and of Definite Proportions . . . . .	21
Nitrogen . . . . .	22
Volumetric Composition of Air . . . . .	23
Water . . . . .	23
Hydrogen . . . . .	26
Composition of Water by Volume and by Weight . . . . .	32
Ozone . . . . .	36
Hydrogen Dioxide and the Law of Multiple Proportions . . . . .	37
Ammonia . . . . .	39
Nitric Acid . . . . .	40
Nitrous Oxide . . . . .	43
Nitric Oxide . . . . .	44
Volumetric Composition of Nitrous and Nitric Oxides and the Law of Volumes . . . . .	46
Chlorine . . . . .	47
Hydrochloric Acid . . . . .	49
Acids, Bases, and Salts . . . . .	52
Carbon . . . . .	54
Marsh Gas . . . . .	55
Ethylene . . . . .	56
Acetylene . . . . .	56
Carbon Dioxide . . . . .	57
Carbon Monoxide . . . . .	59
Influence of Temperature upon the Rapidity of Chemical Action . . . . .	60
Illuminating Gas . . . . .	61
The Blowpipe, Oxidation, and Reduction . . . . .	61
The Equivalent Weights of the Elements, and the Law of Recip- rocal Proportions . . . . .	62

	PAGE
Molecular and Atomic Weights . . . . .	63
Bromine . . . . .	64
Hydrobromic Acid . . . . .	64
Preparation of Ammonium Bromide . . . . .	65
Preparation of Potassium Bromide . . . . .	66
Iodine . . . . .	66
Hydriodic Acid . . . . .	67
Fluorine . . . . .	68
Sulphur . . . . .	69
Hydrogen Sulphide . . . . .	70
Sulphur Dioxide . . . . .	71
Sulphuric Acid . . . . .	73
Phosphorus . . . . .	75
Arsenic . . . . .	77
Antimony . . . . .	79
Boron . . . . .	80
Silicon . . . . .	81
Classification of the Elements and the Periodic Law . . . . .	83
Potassium . . . . .	83
Preparation of Potassium Iodide . . . . .	86
Preparation of Potassium Chlorate . . . . .	87
Sodium . . . . .	87
The Spectroscope . . . . .	89
Ammonium Salts . . . . .	89
Volumetric Composition of Ammonia . . . . .	91
Calcium . . . . .	91
Determination of Carbon Dioxide in Calcium Carbonate . . . . .	93
Barium and Strontium . . . . .	94
Magnesium . . . . .	95
Zinc . . . . .	96
Cadmium . . . . .	97
Mercury . . . . .	97
Copper . . . . .	99
Silver . . . . .	100
Aluminium . . . . .	102
Tin . . . . .	104
Lead . . . . .	105
The Law of Specific Heats . . . . .	106
Bismuth . . . . .	106
Chromium . . . . .	107
Manganese . . . . .	109
Iron . . . . .	111
Nickel and Cobalt . . . . .	113
Table of the Elements . . . . .	114
Table of the Weights of Gases . . . . .	115
Table of Tension of Aqueous Vapor . . . . .	116
Index . . . . .	117

# DIRECTIONS FOR LABORATORY WORK IN CHEMISTRY.



## EXERCISES IN THE CONSTRUCTION OF APPARATUS, ETC.

1. **A Wash Bottle.** — Examine carefully the construction of the sample wash bottle. Soften a cork by means of the cork press, or by rolling it under the foot, and fit it airtight into the neck of a flask of about one liter capacity. Cut the glass tubing to the proper length with the triangular file, and bend the pieces by means of the broad-top burner. Round the ends of each tube by holding it for a short time in the flame. Both in bending and rounding the glass, always keep turning it while in the flame so that it will heat uniformly. Select a cork borer slightly smaller in diameter than the glass tubing, and carefully bore two parallel holes through the cork. Enlarge and smooth the holes with the round file, and fit in tightly the glass tubes. Contract the outside opening of the glass tube which extends to the bottom of the bottle. This can be done either by drawing out the end to a small diameter and cutting it off with the file, or by rounding the end in the flame until the opening is not larger than the diameter of a pin. Clean out the flask and tubes thoroughly, rinse with distilled water, and fill the flask with distilled water.

**2. Glass Stirring Rods.**—Cut a glass rod into short pieces that are of the proper length for stirring liquids in the beakers with which you are provided, and round both ends of each piece in the flame of the burner.

**3. Preparation of Cut Filters.**—Filter paper in sheets, a piece of cardboard, and a pair of compasses are required. A filter when folded and fitted into the funnel must not quite reach the edge of the funnel. Make circles on the cardboard, the radii of which are equal to the length of the sides of the different funnels. Cut out these circles and use them as patterns in cutting the sheets of filter paper.

**4. Folding Filters.**—Practice folding the cut filters (see Fig. 1) and fitting them into the funnels. After having succeeded with the plain filter, inquire of the instructor how star filters are made. Practice pouring water into the filter from a beaker by means of a glass rod.

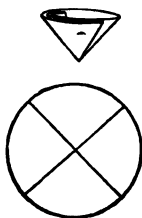


FIG. 1.

**5. Exercise in Manipulation. Crystallization.**—Pulverize a spoonful of crystals of lead nitrate with the mortar and pestle, and pour the powder on a piece of paper. In a beaker, heat water to boiling, and pour into it, from time to time, some of the powder. After each addition of the powder, stir with a glass rod until all is dissolved. Filter the hot solution and evaporate the clear filtrate in an evaporating dish. Set aside the concentrated solution to crystallize. Repeat the experiment, using copper sulphate and potassium bichromate instead of lead nitrate.

**6. Filtration and Washing of Precipitates.**—To a clear solution of lead nitrate add one of potassium bichromate ;



mix thoroughly with a glass rod. A yellow insoluble precipitate, called chrome yellow, is formed. Let the mixture stand for a few minutes until the precipitate has settled; then, by means of the glass rod, carefully pour the clear liquid into a filter. When nearly all has been decanted, add distilled water to the precipitate and stir with the rod. Now allow the precipitate to subside, and again decant the clear liquid. Repeat this washing with water two or three times, or until a drop of the filtrate evaporated on a watch glass leaves no residue. Transfer the precipitate to the filter by means of a jet of water from the wash bottle, allow it to drain, and set it aside to dry.

### CHANGES IN MATTER.

7. By means of pincers hold a piece of platinum wire or foil in the blue flame of the Bunsen burner. Notice how the platinum is changed. Take it out of the flame. Has the platinum been changed permanently? [Write descriptions of the experiments and answers to the questions in your laboratory notebook.]

8. Heat a short piece (two or three inches long) of magnesium ribbon in the same way in which you heated the platinum wire. What happens in this case? Has the magnesium been changed permanently? Has a new substance been formed?

9. Into a clean, dry test tube put enough white sugar to make a layer  $\frac{1}{4}$  of an inch thick. Hold the tube in the Bunsen flame. What changes take place? What do you notice on the sides of the tube? What remains behind? What is its color and taste? Does it dissolve in water?

Is it sugar? Has the sugar been changed permanently into other substances?

10. Dissolve a small quantity (two or three grams) of common salt in water. Evaporate the solution to dryness in an evaporating dish. Examine the residue. Has the salt been changed permanently when it has gone into solution?

11. Into a small tube of hard glass (ignition tube) put enough red oxide of mercury to form a layer  $\frac{1}{8}$  of an inch in thickness. Heat the tube. What change in color do you notice? What is deposited on the sides of the tube? Insert a piece of wood with a spark on the end into the tube; keep on heating. What follows? Take it out and put it back a few times. Is there any difference between the burning in the tube and out of it? What difference? Keep on heating until nothing is left on the bottom of the tube. How do you know that the red substance that you put into the tube has been changed? Has there been a material alteration in composition?

12. Boil water in a beaker. Bring a cold surface into the steam. What collects on the cold surface? Is there any change of composition when water is converted into steam?

In which of the preceding experiments have the substances been changed permanently so that new substances having different properties were formed? In which have the changes been temporary? What are changes of the first kind called? What are those of the second kind called?

Read in your text-book about the difference between chemical and physical changes.

In three of the preceding experiments it was necessary to heat the substances in order that the chemical changes might take place. Under certain conditions light and electricity bring about chemical action. Chemical action, too, can result in the production of heat, light, and electricity. Frequently chemical changes take place by merely bringing substances into contact with one another.

13. Examine a piece of calc-spar or marble. Notice whether it is hard or soft. Heat a small piece in an ignition tube. Does it melt or change in any way? Is the calc-spar soluble in water? In order to learn whether a substance is soluble in water proceed as follows: Put a small quantity of the powdered substance in a test tube with distilled water. Shake thoroughly, and then, as heating usually aids solution, boil. Filter, and evaporate a few drops of the clear filtrate in an evaporating dish or watch glass. If there is anything solid in solution, there will be something solid left in the dish or the watch glass. If not, there will be nothing left. Knowing now the general properties of the calc-spar, you will be able to determine whether it is changed or not. Treat a small piece in a test tube with dilute hydrochloric acid. What takes place? After the action has continued for about a minute, insert a lighted match in the upper part of the tube. Does the match continue to burn? Does the gas in the tube burn? Is the invisible substance in the upper part of tube ordinary air? Why not? Does the calc-spar disappear? In order to tell whether it has changed chemically, the solution must be evaporated. Pour it into a clean evaporating dish and heat with a small flame. Heat gently at the end, and stir with a glass rod to avoid spattering. Examine

the dry substance left in the dish and compare its properties with those of the substance which was put into the test tube. Is it the same substance? Is it hard or soft? Does it change when heated in a tube? Is there an appearance of bubbling when hydrochloric acid is poured upon it? Does it change when allowed to lie in contact with the air? When calc-spar or marble is crushed and ground to a fine powder, is it changed chemically?

What does this experiment illustrate?

14. Bring together in a test tube a small strip of copper and some strong nitric acid. Hold the mouth of the tube away from your face and do not inhale the vapors. What is the appearance of the gas given off? What is the color of the liquid in the tube? Does the copper dissolve? Pour the solution into an evaporating dish and evaporate to dryness. (Conduct this evaporation in the hood, as the vapors of nitric acid are annoying in the air of the room.) In evaporating to dryness be careful not to heat the residue to too high a temperature. Describe the appearance of the substance into which the copper has been changed. Is it soluble in water? What is the effect of heat upon it? When copper is heated so that it melts, is the change chemical?

15. Try the action of dilute sulphuric acid on a little zinc in a test tube. An invisible gas will be given off. Apply a light to the mouth of the tube. What takes place? Add more zinc to the acid, and when it no longer acts on the metal, filter some of the solution and evaporate it to dryness. Carefully compare the properties of the substance left behind with those of the zinc. When an electric current is passed through a strip of zinc, and it acquires new properties, is it a chemical or physical change?

16. Pour concentrated nitric acid upon a bit of tin in a test tube. If no change takes place at first, heat gently, and presently you will have evidence that change is taking place. Is there anything in this experiment which suggests Experiment 14? What is left behind after the action is finished?

What does this experiment show?

As a rule, chemical changes take place more readily in gas mixtures and in solutions than between solids.

17. Mix together in a dry mortar a little dry tartaric acid and about an equal quantity of dry sodium bicarbonate. Do you see any evidence of action? Now dissolve a little tartaric acid in water in a test tube, and a little sodium bicarbonate in water in another test tube. Pour the two solutions together. What evidence have you now that action is taking place? Pour water upon the dry mixture first made. Does action take place? What causes the bubbling? Will a match burn in the gas? In which experiment already performed was a similar gas obtained?

In which of the previous experiments have substances been dissolved in liquids without undergoing chemical change? In which has solution been accompanied by chemical action?

### MATTER AND ENERGY.

Read in the text-books on physics and chemistry that are accessible to you about the relations of matter and energy. Can matter of itself undergo change? What is it that causes changes in matter? What are the different kinds or forms of energy? What is the characteristic property of energy? Do changes in energy accompany changes in composition of bodies? Give illustrations.

**MECHANICAL MIXTURES AND CHEMICAL COMPOUNDS.**

18. Mix two or three grams of powdered roll sulphur and an equal weight of very fine iron filings in a small mortar. Examine a little of the mixture with a strong lens or microscope. Can you distinguish the particles of iron and of sulphur? Pass a small magnet over the mixture. Are particles of iron drawn out of the mixture? Has chemical action taken place? Put some of the mixture in a dry test tube, and heat it with the flame of the burner. Heat gently at first and notice the changes. At first the sulphur melts and becomes dark colored, but soon the whole mass begins to glow, and if you at once take the tube out of the flame, the mass will continue to glow, becoming brighter. Now allow it to cool down and then break the tube and put the contents in a mortar. Does the mass look like the mixture of sulphur and iron with which you started? Examine with the lens and with the magnet. Is the substance now a mixture or a chemical compound? Why?

19. Mix together in a mortar, as intimately as possible, three grams of zinc dust and  $1\frac{1}{2}$  grams of flowers of sulphur. Examine the mixture with a lens of sufficient power to show the grains of sulphur lying side by side with the grains of zinc. To a portion in a test tube add dilute hydrochloric acid. Notice what happens. Make with the rest of the mixture a conical pile on a piece of iron or bit of asbestos paper. Apply the flame of a match, being careful not to stand too near the mixture. What evidence is there that chemical change takes place? Examine with the lens the powder which is left, to see whether there is zinc or sulphur in it. Test some of it with hydrochloric acid. How does it differ from the original mixture?

20. Examine a piece of granite. Is it a mechanical mixture or a compound?

Write in your notebook the definitions for mechanical mixtures and for chemical compounds.

### HOMOGENEOUS MATTER.

Are heterogeneous bodies, such as mechanical mixtures, regarded as distinct substances in chemistry? What do we mean by the term *homogeneous substance*?

Make a list of homogeneous and heterogeneous substances in your notebook.

Read in your lecture notes and text-books in regard to the meaning of the terms *element*, *compound*, and *physical mixture*. Give examples of each class.

Would it be possible from the appearance of a homogeneous substance to tell to which of these classes it belonged? How would you determine whether a given homogeneous substance was an element, a compound, or a physical mixture?

State approximately how many elements are known at present.

Read in your text-book about the relative abundance of the elements.

Which is the most abundant element?

Which elements make up three fourths of the solid crust of the earth?

Which elements are the chief constituents of plants and animals?

Look over the list of elements on page 114, and notice how the symbols have been formed.

In what respects do physical mixtures differ from mechanical mixtures?

### THE CHEMICAL BALANCE.

**21.** Examine the construction of the analytical balance. Find out from the instructor how to use it. Weigh accurately a clean watch crystal. The object to be weighed must always be placed on the left-hand scale pan. Never handle the weights with the fingers, but always use the pincette provided for the purpose. In placing weights on the right-hand scale pan follow the directions of the instructor, who will also explain to you how to adjust the rider and to count up the weights properly.

**22. Determination of Specific Gravity.** — Select a small piece of some solid which is insoluble in water, — for example, a piece of marble or a bit of iron or brass. Attach a short piece of thread to it so that it can afterwards be suspended in water contained in a small beaker. Weigh it carefully in the air. Then suspend it, by means of the thread, from the hook above the scale pan so that it is submerged in water in a small beaker. Determine its weight in the water. Calculate the specific gravity.

Determine also the specific gravity of a liquid, such as alcohol, using a small specific gravity bottle. First weigh the clean, dry bottle, then weigh it filled with alcohol. Wash out the alcohol with water, and now weigh it filled with water. Calculate the specific gravity.

### MEASUREMENT OF LIQUIDS.

**23.** For measuring liquids the liter and its subdivisions are used. Find out the relation between the unit of weight and the unit of capacity in the metric system. How is the liter related to the meter ?



Compare the liter flask with the half- and quarter-liter flasks. Inquire of the instructor how to use the pipettes, burettes, and graduated cylinders. Measure the capacity of two large glass bottles. In measuring liquids be careful that your eye is in the same plane with the surface.

**24.** Construct a graduated test tube by running in accurately 5 or 10 cubic centimeters of water from a burette. Mark the height at which the water stands with a strip of adhesive paper. Then add another 10 cubic centimeters, mark as before, and so continue until within half an inch of the top. Afterwards make permanent marks on the glass with a file or a diamond.

### THE AIR.

#### [LABORATORY DEMONSTRATION.]

**25. Air has Weight.** — To determine roughly the approximate weight of air, select two flasks of about 250 cubic centimeters' capacity which have been blown in the same mold and are therefore of equal size. Fit them tightly with corks. Cork one permanently and reserve it as a counterpoise. Add to the other about 25 cubic centimeters of water, and boil the water over a lamp until the interior is full of steam and all air has been expelled; then quickly cork tightly, remove the lamp, and allow the flask to cool. Tare now the second flask with the first and such additional weight as may be necessary. [In weighing the flasks, use a large lecture table balance.] Draw the cork of the flask from which the air has been expelled, and after the air has entered, determine the increase in weight. Determine the volume of the air weighed by adding water

from a graduated measure until the flask is filled to the former level of the cork. From these values the weight of a liter of air under the conditions of the experiment can be calculated. Read in a text-book on physics how the accurate value for the weight of a liter of air has been determined.

**26.** The pressure of the atmosphere is measured by means of the barometer. Describe in your notebook how a barometer is made. Read the barometer in the laboratory, and enter the value in your notebook, together with the date and time of day.

**27.** In a small clay or porcelain crucible put a small piece of zinc. Support the crucible on a pipe-stem triangle, and heat it with the burner. After the zinc is melted, stir it with an iron wire. Continue to heat and stir until it is no longer liquid. Describe what has taken place.

Now repeat the experiment, but before heating the zinc, add enough dry borax powder to form a complete cover over the metal when both are melted. It may be necessary to heat the crucible with the blast lamp. Is the zinc changed to a powder in this case? What conclusion can you draw from this experiment?

**28. Metals when Heated or when Burnt in the Air Increase in Weight.** — Place a saltspoonful of pulverized iron in a small crucible, then put in a short piece of iron wire that will serve as a stirring rod, and now weigh the crucible with the prescription scales. Note the weight in your laboratory book. Heat the crucible with the burner, stir the powder with the wire, taking care not to lose any of the powder. Reweigh when cold, and describe in your notes what has happened.

29. Repeat the preceding experiment, using about  $\frac{1}{2}$  of a gram of magnesium powder instead of the iron.

[LABORATORY DEMONSTRATION.]

30. **A Candle and Other Combustibles in Burning Increase in Weight.** — On one pan of the prescription scales place a short candle, and directly above it suspend a wide glass tube or lamp chimney containing pieces of caustic soda, a substance that has the power of absorbing the products of combustion of the candle. See Fig. 2. Bring the appara-

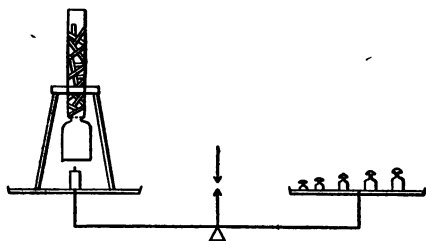


FIG. 2.

tus into equilibrium with weights; then light the candle and let it burn for a few minutes. When the apparatus has become cold, it will be found to be heavier than before the candle was burnt. What inference can be drawn from this and the two preceding experiments?

31. **Burning Substances only Combine with a Portion of the Air.** — Float an evaporating dish on water in the pneumatic trough. Put a piece of phosphorus a little larger than a pea in the dish. Observe great care in handling the phosphorus, using the forceps and cutting it under the water. Set fire to the phosphorus with a hot wire, and quickly place a glass-stoppered bell jar over the burning

phosphorus, and let the bell jar dip an inch or two into the water. At first some air will be driven out of the bell jar on account of the expansion due to heat. After the



FIG. 3.

burning has stopped, allow it to stand for some time. The fumes gradually disappear, and an invisible gas remains. Notice that the water has risen in the bell jar. Bring the bell jar into such a position that the height of the water is the same on the outside and inside. Take out the stopper and insert a burning splinter

or a burning candle. Does it continue to burn? What conclusions do you draw from the results of this experiment in regard to the composition of the air?

### MEASUREMENT OF GASES.

32. Gases are measured in eudiometers and in gas measuring tubes. Notice how these are constructed. Thermometers, a barometer, and pneumatic troughs are also required.

To compare one gas volume with another, it is necessary that they should be at the same temperature and under the same pressure. Why is this so?

What effect does increase of temperature have upon the volume of a gas?

What is the law of Charles? (Gay-Lussac's Law?)

What effect do changes of pressure have upon the volume of a gas?

What is Boyle's law? (Mariotte's Law?)

The standard temperature and pressure for comparing gas volumes is 0° Centigrade and 760 millimeters of mercury pressure.

**33.** From the laws of Charles and Boyle, deduce the following formula for reducing gas volumes to standard conditions:—

$$V' = \frac{VP}{760 \left( 1 + \frac{t}{273} \right)}$$

In this formula

$V'$  = the volume of the gas under standard conditions.

$V$  = the volume of the gas under a pressure of  $P$  millimeters of mercury and at a temperature of  $t$  degrees.

**34.** Measure a volume of air in a eudiometer over mercury. Read the temperature by means of a thermometer that dips into the mercury. Read the height of the barometer. Notice whether the mercury is at the same level on the outside and inside of the eudiometer. How does this difference of level affect the pressure upon the gas? Measure the difference of level and determine the pressure of the gas. Now calculate what the volume would be under standard conditions.

The usual method of determining the weight of a volume of gas is to calculate what the volume would be under standard conditions, and then to multiply by the weight of one volume under standard conditions. See table of the weights of gases, page 115.

A gas that is saturated with moisture occupies a larger volume or exerts a greater pressure than it would if it were perfectly dry. It is really a mixture of the dry gas with the vapor of water. To reduce such a gas volume to standard conditions, it is necessary to subtract from the apparent

pressure of the gas the pressure which the aqueous vapor exerts at the observed temperature. Then proceed as in the previous case. See Table of Tension of Aqueous Vapor, page 116.

### OXYGEN, O.

**35. Preparation of Oxygen from Mercuric Oxide.** — Place in an ignition tube a layer of mercuric oxide about  $\frac{1}{4}$  of an inch in thickness. Close the tube with a stopper provided with a delivery tube. Support the tube with the clamp

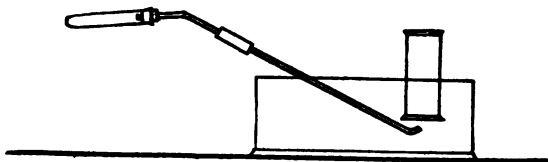


FIG. 4.

stand in the position shown in Fig. 4, so that the delivery tube is under the cylinder filled with water in the pneumatic trough. Heat the ignition tube, and collect the gas in gas bottles. What is formed on the sides of the ignition tube? When the evolution of gas has nearly stopped, take the delivery tube out of the water, then take away the flame from the ignition tube. Examine the gas by inserting a stick with a spark on the end. What takes place? Is the gas contained in the vessel ordinary air?

**36.** Repeat the preceding experiment, using potassium chlorate instead of mercuric oxide.

**37.** Weigh off 10 grams of potassium chlorate, and grind it in the mortar to a fine powder; then mix with it intimately 10 grams of powdered manganese dioxide. Fill the ignition tube one half full of the mixture, heat, and fill all your gas bottles with oxygen.

**38. Physical Properties of Oxygen.** — The gas in the bottles is invisible. The slight cloud which appears when the gas is first collected, is due to the presence of a small quantity of solid particles that have been carried over by the gas. After standing a short time this cloud disappears. The gas is tasteless and inodorous. (Inhale a little from one of the bottles.) It is slightly heavier than the air. When subjected to a sufficiently low temperature and high pressure it becomes liquid.

**39. Chemical Properties of Oxygen.** — Into a jar of oxygen introduce a little lump of roll sulphur by means of a deflagrating spoon. Let it stand a short time, and notice what change, if any, takes place. Repeat this experiment, using charcoal and phosphorus in place of the sulphur. [Phosphorus should be handled with great care. It is always kept under water. If a small piece is wanted, take out a stick with the forceps and put it under water in an evaporating dish or pneumatic trough. *While it is under water* cut off a piece of the size wanted. Take this out by means of the forceps, lay it for a moment on a piece of filter paper to absorb the water, then quickly put it in a deflagrating spoon.] What does this experiment show?

**40.** In a deflagrating spoon set fire to a little sulphur and let it burn a short time in the air. Notice whether it burns with ease or with difficulty. Notice the odor of the fumes. Now plunge the burning sulphur into a bottle of oxygen (Fig. 5). Does it burn more readily in oxygen than in the air? Notice the odor of the fumes in the bottle. Is it the same as that noticed when the burning took place in the air?



FIG. 5.

41. Fasten a bit of charcoal to a wire, heat it in the flame until it is red hot, then plunge it into oxygen. How does the burning in oxygen compare with the burning in air?

42. Burn a *small* piece of phosphorus in the air and in oxygen. In the latter case the light emitted from the burning phosphorus is so intense that it is painful to some eyes to look at it. After the burning is over let the bottle stand. Does it become clear?

43. Insert a stick with a spark on the end into the bottles in which sulphur, charcoal, and phosphorus have been burnt. Does it continue to burn? Is there any oxygen left? What do you infer from this?

44. Straighten a steel watch spring by taking hold of the ends and slowly passing it through the flame until all parts have been red hot. Wind a little thread around one end and dip it into melted sulphur. Set fire to this and plunge it into a vessel containing oxygen. For a moment the sulphur will burn, but soon the steel will begin to burn brilliantly, and the burning will continue as long as there is oxygen left in the vessel. The phenomenon is of great beauty, especially if observed in a dark room. The walls of the vessel become covered with a dark reddish-brown substance. Some of the products of combustion will also be found at the bottom in the form of globules.

45. Into another jar of oxygen plunge a lighted strip of magnesium ribbon. Hold the magnesium with the forceps, and use a piece about four inches long.



**46.** Fasten a short piece of a candle to a bent wire so that it can be lowered into a gas bottle (Fig. 6). Light the candle and insert into a jar of oxygen. Notice the appearance of the flame. How does the burning in oxygen compare with the burning in air? Keep the jar in which the candle is burning covered, and notice what happens after a time. Why is the candle extinguished?



FIG. 6.

Write in your notebook a full account of all the experiments you have made with oxygen.

What inference can you draw from these experiments in regard to the nature of combustion?

State the characteristic physical and chemical properties of oxygen.

Read in the text-books about the slow action of oxygen at ordinary temperature. Give examples of substances that are acted upon by oxygen at ordinary temperature.

[LABORATORY DEMONSTRATION.]

**47. Experiment to show that Oxygen Disappears when Phosphorus Burns in it.** — Float a small dry evaporating dish on the water of the pneumatic trough. Put a dry piece of phosphorus as large as a bean in the dish. Place over it a glass-stoppered bell jar of two liters' capacity. Remove the stopper and sink the bell jar so far into the water that it is half full of water and the level is the same outside and inside. Fasten the bell jar securely in this position, and now displace the air in it with oxygen. This can be done by depressing the oxygen delivery tube until it is under the edge of the bell jar and then allowing a slow current of the gas to enter. After about two liters of oxygen have been passed in, the stopper having been

removed during this time, it may be assumed that the air has all been displaced. Now heat the end of a piece of wire red hot, plunge it into the bell jar and set fire to the phosphorus in the dish. Quickly withdraw the wire and instantly close the bell jar tightly with the stopper. At first the heat expands the gas, but it soon begins to contract, and finally, if enough phosphorus was present, and all the air was displaced, the water will completely fill the vessel.

What has become of the oxygen? What would have happened if under the same conditions sulphur and charcoal had been burnt in oxygen? If iron or magnesium had been used in place of phosphorus? If the phosphorus had been burnt in air contained in the bell jar instead of in oxygen? In which previous experiment have you tried this, and what was the result? From all the experiments with oxygen, what do you conclude in regard to the nature of combustions in air and in oxygen?

### THE CONSERVATION OF MATTER.

48. Select two small beakers and fill one of them about one third full of a solution of barium chloride; fill the other one third full of dilute sulphuric acid. Place both beakers in the scale pan of a prescription balance, and bring it into equilibrium. Now carefully pour one solution into the other; do not lose a drop, and place the empty beaker upon the scale pan. Has a chemical change taken place? Has there been any change in weight?

What is the law of the conservation of matter?

Can energy be created or destroyed? Give reasons for your answer. State the law of the conservation of energy. What does it mean?

## LAWS OF CONSERVATION OF MATTER AND DEFINITE PROPORTIONS.

### 49. Analysis of a Known Weight of Potassium Chlorate.

—Select a clean piece of hard glass tubing about five inches long and not more than  $\frac{1}{4}$  of an inch in outside diameter. Draw out one end in the flame and seal it. Round the edges of the other end in the flame. When cold, weigh the empty tube with the analytical balance. Note the weight in your laboratory notebook. Put into the tube a few small crystals of potassium chlorate. [Less than  $\frac{2}{10}$  of a gram should be taken.] Weigh again. The difference between the second and first weighings is the amount of chlorate in the tube. Push into the middle of the tube a small plug of glass wool or asbestos fiber. Weigh again, and then bend the tube as shown in the figure. When the tube is cold, connect it by means of a capillary tube and two rubber tubes with the gas measuring burette as shown in Fig. 7. The rubber tube that connects with the burette should be provided with a pinchcock, and the burette should be full of water to the 0 mark. Now open the pinchcock and notice whether the joints are air tight. If they are, the water in the burette will not fall, although the water in the pressure tube stands but an inch or two above the base. If the joints are air tight, heat the crystals gently with a small flame, and continue heating until no more gas is evolved. Let the apparatus stand for half an hour, or until it has acquired the

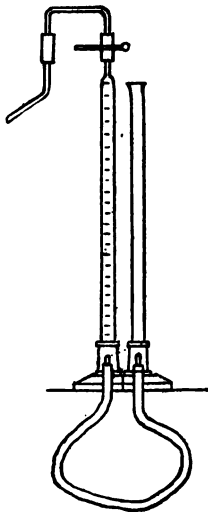
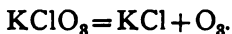


FIG. 7.

temperature of the room. Read the height of the barometer; take the temperature of the gas by means of a thermometer attached to the outside of the tube. Bring the level of the water in the burette and the pressure tube to the same height and read the volume of gas. From the observed volume, temperature, and pressure calculate the weight of oxygen that has been obtained from the known weight of potassium chlorate. Disconnect the tube containing the potassium chloride and weigh it. Calculate the weight of potassium chloride that is left. The weight of oxygen found added to the weight of potassium chloride should equal the weight of potassium chlorate taken.



A repetition of the experiment with a different quantity of potassium chlorate will show that this compound has a definite and invariable composition.

What is the law of definite proportions? Find out from your instructor and your text-books, how, from the laws of conservation of matter and definite proportions, chemical changes can be represented by means of equations like the one above.

### NITROGEN, N.

50. Nitrogen is most conveniently prepared by burning phosphorus in air confined in a bell jar over water. The phosphorus forms a white, solid oxide which readily dissolves in the water. Prepare nitrogen in this way, using the apparatus described in Experiment 31, under The Air. What is the color, odor, and taste of the gas? Try the effect of introducing into it, successively, several burning bodies, as, for example, a candle, a piece of sulphur, phosphorus, etc. Do these substances continue to burn?

**VOLUMETRIC COMPOSITION OF THE AIR.**

**51.** A gas burette, such as was used in the analysis of potassium chlorate, and a phosphorus absorption pipette are required. One hundred cubic centimeters of air under atmospheric pressure and temperature are taken into the burette and transferred to the absorption pipette. The phosphorus oxidizes slowly, and the air is filled with white fumes. After some minutes these disappear. Now draw back the residual nitrogen and measure the volume. Again pass the gas into the pipette, let it remain a short time, and then bring it back into the burette. If no further contraction has taken place, all the oxygen has been removed.

Repeat the analysis until good results are obtained. Is the air a chemical compound, or a physical mixture? What are your reasons for regarding it as such?

**WATER,  $H_2O$ .**

**52. Occurrence and Distribution.** — Read in your text-book about the occurrence of water. Many apparently dry substances contain water. Put into a dry test tube a small piece of wood, and heat gently. What evidence do you obtain that water is given off? Do the same thing with a piece of fresh meat.

**53. Water of Crystallization.** — Heat gently a few crystals of alum in a test tube. Describe what takes place.

Perform a similar experiment with some gypsum, which is the natural substance from which plaster of Paris is made.

Heat also in a tube a few small crystals of copper sulphate, or blue vitriol. In this case the loss of water is

accompanied by the loss of color. After all the water is driven off, the powder left behind is white. On dissolving it in water, however, the solution will be seen to be blue, and if the solution is evaporated, blue crystals will be again obtained.

**54. Efflorescent Substances.** — Select a few clear, transparent crystals of sodium sulphate, or Glauber's salt. Put them in a watch glass, and let them lie exposed to the air for an hour or two. They soon lose their luster, and become white and powdery on the surface. What is the cause of this?

**55. Deliquescent Substances.** — Expose a few pieces of calcium chloride to the air. Its surface will soon give evidence of being moist, and after a time the substance will dissolve in the water which is absorbed.

**56. Physical Properties of Water. Melting and Boiling Points.** — For this experiment a thermometer with a scale from  $-5^{\circ}$  to above  $100^{\circ}$  is required. Place some broken ice in a beaker and insert the thermometer; let it remain for some time, then read the temperature. Does the quantity of ice melting make any difference in the temperature as indicated by the thermometer?

Boil water in a beaker, and take the temperature with the thermometer. Does the quantity of water that is boiled or the rapidity of boiling make any difference in the temperature?

How does the temperature of the ice compare with that of the lambent water, or the temperature of boiling water with that of the steam above it? What becomes of the heat that enters the containing vessel?

What is the effect of a change of atmospheric pressure

upon the boiling point? What effect do changes of pressure have upon the melting point?

What is the specific gravity of ice?

**57. Distillation of Water.** — Use for the purpose a glass retort holding about 200 cubic centimeters. Dissolve a few crystals of copper sulphate in 100 cubic centimeters of water. Pour the solution into the retort. Fasten the latter in the retort stand. Over the neck of the retort push the neck of a glass flask, and allow the flask to dip in cold water contained in the pneumatic trough. See Fig. 8. Distill the solution until only a little is left in the retort. What becomes of the copper sulphate? What would happen if a volatile substance had been dissolved in the water instead of the non-volatile copper sulphate?

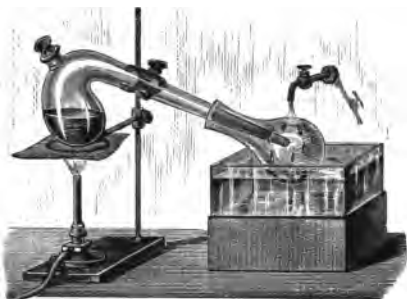


FIG. 8.

[LABORATORY DEMONSTRATION.]

**58. Decomposition of Water by Means of an Electric Current.** — Connect the wires from the poles of a Bunsen battery with the platinum electrodes of a U-tube apparatus for decomposing water. The apparatus is filled with water, to which has been added one tenth its volume of sulphuric acid. Start the current, and gas bubbles will collect in each of the two branches of the tube. After the current has been going for a few minutes, notice that there is twice as much gas in one branch as in the other. When one

branch is full of gas, stop the current. Hold a lighted match above the stopcock of the branch containing the larger gas volume. Open the stopcock and light the escaping gas. A flame will be noticed. The gas burns. Is it air, or oxygen, or nitrogen? Bring a splinter with a glowing end over the other stopcock and allow the gas to escape. Does the gas act like oxygen? Describe this experiment in your notebook.

### HYDROGEN, H.

**59. Preparation of Hydrogen.** — Throw a *small* piece of sodium<sup>1</sup> on water. While it is floating on the surface apply a lighted match to it. What takes place? The flame is due to burning hydrogen. The yellow color of the flame is caused by the presence of sodium, some of which also burns. If a piece of filter paper is floated on the water, and the sodium is then thrown upon the moist paper, the hydrogen will take fire spontaneously. This is also the case if the sodium is thrown upon warm water.

A small piece of potassium thrown upon water decomposes it in the same way; the hydrogen takes fire even when the metal is thrown upon cold water, and burns with a violet flame. The color is due to the presence of potassium, some of which also burns. Fill a test tube with water in the pneumatic trough, and, by means of a sharp-pointed file, pick up a small piece of sodium and quickly plunge it under the opening of the tube. The sodium will ascend to the top, and in this way the gas can be

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<sup>1</sup> Sodium and potassium must be handled with care. It is not advisable to use a piece larger than a grain of wheat. Sodium and potassium are preserved under oil. Phosphorus is kept under water. Do not make the mistake of trying to cut sodium or potassium under water.



collected. Add more sodium until the tube is full of gas. Put your thumb over the mouth of the tube, bring it out of the water, and test the gas with a lighted match. Does it burn?

Test the water in the trough with red litmus paper. What change do you notice?

Explain the chemical changes that have taken place.

[LABORATORY DEMONSTRATION.]

**60.** Certain metals which do not appreciably decompose water at ordinary temperature, or which act upon it slowly, decompose it easily at elevated temperatures. This is true of iron and zinc. If steam is passed through a tube con-

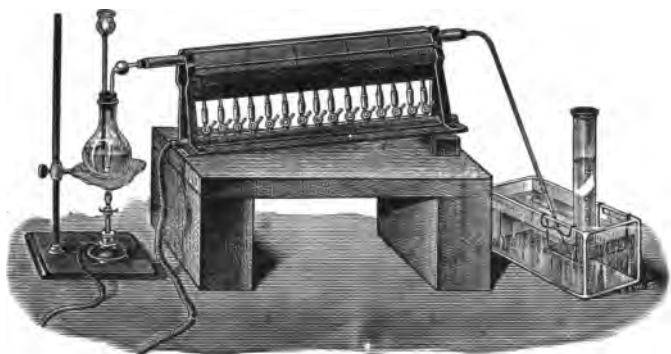


FIG. 9.

taining iron turnings or fine, bright iron wire nails heated to redness, the water is decomposed, the oxygen is retained by the iron in chemical combination, while hydrogen is liberated. Zinc dust decomposes water at  $100^{\circ}$ . Fill a hard glass combustion tube with zinc dust, and pass a current of steam through the tube. Gently heat the tube in a furnace, and collect the hydrogen over the

water in the pneumatic trough. See Fig. 9. Light the hydrogen and observe the color of the flame.

Magnesium powder, when heated in a current of steam, burns brilliantly; hydrogen is given off and may be collected as in the case of the zinc.

**61. Hydrogen prepared from Acids.** — Into a test tube, put a few pieces of granulated zinc, and pour upon the zinc enough hydrochloric acid to cover it. What do you notice? After the action has continued for a minute or two, apply a lighted match to the mouth of the tube. What takes place? Try the same experiment with dilute sulphuric acid. What is the result? The gas given off is hydrogen. For the purpose of collecting the gas the operation is best performed in a flask of about 250 cubic centimeters' capacity. Fit a rubber stopper with two holes into the neck of the



FIG. 10.

flask. Into one of these holes put a funnel tube, which must reach nearly to the bottom of the flask. Into the other put a glass tube bent at right angles. This tube need only pass through the stopper and not extend below it. The outside end of this tube is connected by means of a short piece of rubber tubing,

with a delivery tube which dips under the water in the pneumatic trough. [Instead of the bottle shown in the figure an ordinary flask may be used, provided care is taken not to break it in putting in the zinc.] Put a small handful of granulated zinc into the flask, and pour upon it

enough of a cooled mixture of sulphuric acid and water (one volume of concentrated acid poured into six volumes of water — *never pour water into concentrated sulphuric acid!*) to cover it. Usually a brisk evolution of gas will take place at once. *Wait two or three minutes*, and then collect some of the gas by displacement of water. Should the action become slow, add a little more dilute acid. Fill several cylinders and bottles with the gas.

Explain the chemical change that takes place in the flask. Evaporate some of the liquid in the flask, and obtain zinc sulphate.

**62. Physical Properties of Hydrogen.** — Notice that hydrogen is a colorless, tasteless, inodorous gas. When made by the action of zinc on acids, it has a slightly disagreeable odor. This is due to the presence of small quantities of impurities. If these are removed, the odor disappears. This can be done by passing the gas through a solution of potassium permanganate.

Hydrogen is not poisonous and may be inhaled with impunity. It is the lightest substance known. The air is 14.4 times heavier; oxygen is 16 times heavier. Place a vessel containing hydrogen with the mouth upward and uncovered. In a short time examine the gas in the vessel and see whether it is hydrogen. What has happened?

**63. Hydrogen Poured Upward.** — Hold a glass cylinder, containing air, inverted in one hand; with the other hand bring a cylinder full of hydrogen from the pneumatic trough, and gradually pour the hydrogen up into the inverted cylinder. The air will be displaced. On examination the inverted cylinder will be found to contain hydrogen, while the one with the mouth upward will contain

none. Soap bubbles or small collodion balloons filled with hydrogen rise in the air. Fill a collodion balloon with hydrogen, using the Kipp's apparatus to generate the gas. When full, close the opening by means of a string, and let it go. Large balloons are always filled with hydrogen or some other light gas. Some kinds of illuminating gas are rich in hydrogen, and may be used for the purpose.

What is the weight of a liter of hydrogen under the standard conditions? What is the weight of a liter of oxygen under the same conditions? Divide the latter by the former.

[LABORATORY DEMONSTRATION.]

**64. Diffusive Power of Hydrogen.** — A porous clay battery jar is closed with a cork, and a long glass tube open at both ends is passed through the cork. Fasten the apparatus in a stand, and make the end of the glass tube dip into water contained in a beaker, or arrange it as shown in Fig. 11. Now carefully bring a bell jar full of hydrogen over the battery cup, and notice what happens. Take away the hydrogen vessel, and observe what takes place. Write out a description and explanation of this experiment in your notebook.



FIG. 11.

What is the law of the diffusion of gases?

What are the relative rates of diffusion of hydrogen and oxygen?

**65. Chemical Properties of Hydrogen.**—Hold a wide-mouthed bottle or cylinder filled with hydrogen with the mouth downward. Insert into the vessel a long lighted taper. The gas takes fire at the mouth of the vessel, but the taper is extinguished. On withdrawing the taper and holding the wick for a moment in the burning hydrogen, it will take fire ; but on putting it back in the hydrogen, it will be again extinguished. What does this experiment show? Other burning bodies behave in the same way.

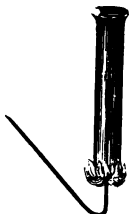


FIG. 12.

**66.** Pass hydrogen from a generating flask, or Kipp's apparatus, through a U-tube containing granular calcium chloride, to dry the gas. Connect by means of rubber tubing the nozzle of a blowpipe with the U-tube. After the hydrogen has been allowed to escape for a few moments, set fire to it. [Always be cautious in working with hydrogen. The danger consists in the fact that a mixture of hydrogen and oxygen, or hydrogen and air, is explosive. It requires a flame or spark to explode it. Always let the gas escape for a time, and collect a test tube full, and light it to see if it will burn quietly before applying a flame to the hydrogen escaping from the apparatus.] In a short time it will be seen that the flame is practically colorless, and gives no light. That it is hot is shown by holding a piece of platinum or a piece of some other metal in it.

Hold a dry vessel over the flame of burning hydrogen. Drops of water will condense on its surface and run down. If a cold, dry bell jar is held above the flame, the steam formed will be condensed and drops of water will collect on the walls of the vessel.

What conclusion do you draw from this experiment in regard to the composition of water?

**67.** A strong glass cylinder is filled one third full of oxygen in the pneumatic trough. The rest of the water is then displaced by hydrogen. Take it out of the water covered with a glass plate. Remove the plate, and at the same instant apply a light. A sharp explosion will occur. What is the cause of the explosion?

**68.** Great heat is evolved when hydrogen and oxygen combine. A jet of oxygen issuing into a flame of burning hydrogen produces an intense heat, and is used practically for melting platinum and for producing the lime light. The apparatus used is called the oxyhydrogen blowpipe. Examine the construction of the oxyhydrogen blowpipe.

[LABORATORY DEMONSTRATION.]

**69.** Hold in the flame of the oxyhydrogen blowpipe successively a piece of iron wire, a steel watch spring, a piece of copper wire, a piece of zinc, a piece of platinum wire.

Allow the flame to strike upon a piece of quicklime. Describe the effects produced.

Make a tabular statement of the characteristic properties, both physical and chemical, of hydrogen.

**COMPOSITION OF WATER BY VOLUME AND BY WEIGHT.**

[LABORATORY DEMONSTRATION.]

**70.** Into the eudiometer filled with mercury and standing in the mercury dish as shown in Fig. 13, bring about 10 cubic centimeters of oxygen. Press the eudiometer

down on the conical stopper and raise the pressure tube until the level of the mercury is the same outside and inside of the eudiometer. Read the volume of oxygen taken. Bring the eudiometer into its original position and now introduce not less than 25 cubic centimeters of hydrogen free from air. Again read the volume of the gas when it is under atmospheric pressure. Now expand the gas mixture very considerably by lowering the pressure tube, and close the rubber tube with a strong pinchcock. Explode the mixture by passing an electric spark through the tube. Read the volume of gas that remains, and calculate what volume of hydrogen has united with the oxygen. Test the residual gas; it will be found to be hydrogen.

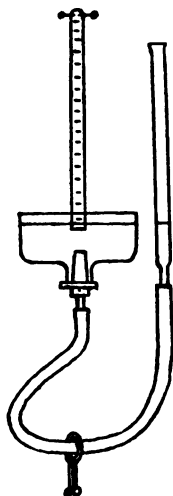


FIG. 13.

Exactly two volumes of hydrogen unite with one volume of oxygen. If either gas is present in excess of this proportion, the excess will remain uncombined.

How many volumes of steam are formed by the union of two volumes of hydrogen with one of oxygen at  $100^{\circ}$ ? How could this be shown by experiment? Knowing the composition of water by volume and the relative densities of hydrogen and oxygen, what is the composition of water by weight?

In Experiment 70, water may be used instead of mercury to fill the eudiometer and the glass dish and pressure tube.

Another apparatus that may be used for this experiment consists of a gas measuring burette and an explosion pipette, as shown in Fig. 14. The burette is filled with

water; then about 10 cubic centimeters of oxygen, free from air, are introduced from a gasometer. The volume of oxygen is read off when it is under atmospheric pres-

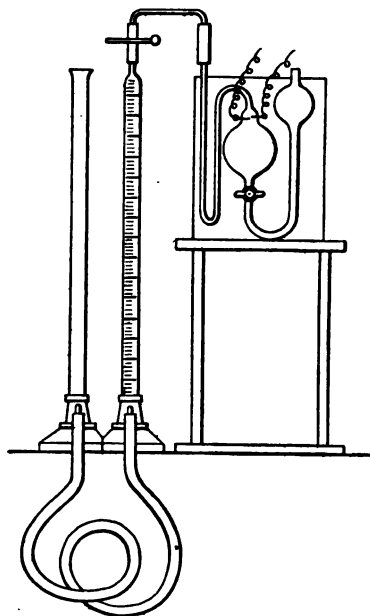


FIG. 14.

sure. About 25 or 30 cubic centimeters of hydrogen are then introduced and the volume is again determined under atmospheric pressure. [In introducing the hydrogen and oxygen into the gas burette it is advisable to use a T-tube in connecting the burette with the gasometer or generating flask. By this means it is possible to drive out all air from the connections before admitting the gas to the burette.] The difference between the first and second readings gives the volume of hydrogen taken. Now trans-



fer the mixed gases to the explosion pipette, and after closing the valves explode the mixture by passing a current of electricity through the wires. Draw back the residual gas into the burette and measure its volume. Test it by means of a lighted match on expelling it into the air, to show that it is hydrogen.

**71. Gravimetric Composition of Water.**—Construct an apparatus like that shown in Fig. 15. A and B are cylinders half filled with concentrated sulphuric acid, C is a tube of

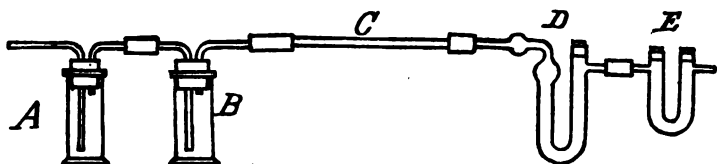


FIG. 15.

hard glass, the middle portion of which is filled with granular copper oxide. The copper oxide is held in place by plugs of glass wool or asbestos fiber. D and E are tubes filled with calcium chloride. C and D must be accurately weighed before the experiment on the analytical balance. A Kipp's hydrogen generating apparatus is now connected with A and a slow current of hydrogen is allowed to drive out all the air in the apparatus. When this has been done, heat the copper oxide in C, gently at first, afterwards more strongly. Notice that the copper oxide glows, and drops of water condense in D. Keep the end of C near to D warm, from the beginning of the experiment until the end, so that water does not condense in it. When all the copper oxide is reduced, take away the flame, and allow the tube to cool in the current of hydrogen. Disconnect C and

D. Draw dry air through them to displace the hydrogen and weigh them. The increase in the weight of D is the weight of water formed. The loss in weight of C is the weight of oxygen contained in the water.

If the experiment is carefully performed, it will be found that the weight of the oxygen is eight ninths the weight of the water.

The tube D can be used several times without refilling. When not in use it is closed with short pieces of rubber tubing plugged with glass rods. E is to prevent moisture of the air from entering D, and A and B serve to dry the hydrogen before it enters the tube C. It is advisable to have the end of the tube C touching the end of D inside of the rubber connection, in order that water may not be condensed on the rubber. The rubber need not be more than  $\frac{1}{2}$  of an inch in length.

Read in one of the larger works on chemistry an account of the method used by Dumas in determining the gravimetric composition of water.

What is meant by "oxidation" and "reduction"? Give illustrations of each kind of action.

### OZONE, $O_3$ .

72. Place a few pieces of phosphorus in a cylinder or gas jar and add enough water to partially cover it. Cover the vessel and allow it to stand for 20 minutes. What evidence is there that the phosphorus is slowly oxidizing? Notice the odor of the air in the flask. Boil water in a small beaker or test tube and add a very small quantity of starch powder. After the starch has dissolved, add a crystal of potassium iodide. Moisten a strip of filter paper with this solution and introduce it into the cylinder con-

taining the phosphorus. What change do you notice? Does the air, or oxygen, or nitrogen act upon potassium iodide and starch? What has caused the chemical change?

**73.** Make a solution of manganese sulphate in water, dip strips of filter paper into it, and insert them into the ozone vessel. What change takes place?

By what other methods can ozone be obtained?

Describe the manner in which ozone can be converted into ordinary oxygen.

#### **HYDROGEN DIOXIDE, $H_2O_2$ , AND THE LAW OF MULTIPLE PROPORTIONS.**

**74.** Pour 10 cubic centimeters of dilute sulphuric acid into 20 cubic centimeters of water, and add gradually 10 grams of barium dioxide. Stir constantly with a glass rod while adding the powder. Filter the solution. The clear filtrate contains hydrogen dioxide. Add a few drops of the solution to a solution of potassium iodide and starch. What effect does it have? To another portion of the hydrogen dioxide solution add a small quantity of powdered manganese dioxide. Notice that gas bubbles escape. What is the cause of this? To another portion add, drop by drop, a very weak solution of potassium permanganate. Does the violet color disappear?

The composition of hydrogen dioxide has been found to be one part of hydrogen to 16 parts of oxygen by weight. How does this compare with the composition of water by weight? How do the properties of hydrogen dioxide compare with those of water?

What is the law of multiple proportions?

## [LABORATORY DEMONSTRATION.]

**75. Experiments to illustrate the Law of Multiple Proportions.** — Clean a small porcelain crucible, support it on a tripod by means of a pipe-stem triangle. Heat it for a few moments with the blue flame of the burner, then while it is still warm, place it by means of pincers in a desiccator containing pieces of calcium chloride. After the crucible has stood in the desiccator for 15 or 20 minutes, weigh it accurately with the analytical balance. Note the weight carefully in your notebook. Cover the bottom of the crucible with a layer of dry copper oxide  $\frac{1}{8}$  of an inch in thickness. Weigh the crucible again. The increase in weight is the weight of copper oxide in the crucible. Place the crucible again on the pipe-stem triangle, and cover it with a lid which has a small hole in the middle. Through this hole a porcelain tube bent at right angles is passed. (Rose's crucible.) The end of the tube is connected with a Kipp's apparatus for hydrogen, and a slow current of this gas is turned on. After a minute or two, when all the air has been expelled from the crucible, the copper oxide is heated in the atmosphere of hydrogen for 8 or 10 minutes. The crucible is then allowed to cool down in the stream of hydrogen. When cold, it is again weighed. The difference between the third and second weighings gives the weight of oxygen in the known weight of copper oxide. Calculate what weight of copper is combined with eight parts by weight of oxygen in copper oxide.

Now repeat the experiment, using red cuprous oxide instead of black cupric oxide, and from the results obtained calculate the weight of copper that is in combination with eight parts of oxygen in this compound. How

do the proportions of copper in the two compounds compare with each other?

Read in the text-books about other instances of chemical combination which illustrate the law of multiple proportions.

### AMMONIA, $\text{NH}_3$ .

76. To a little ammonium chloride on a watch glass add a few drops of a strong solution of caustic soda, and notice the odor of the gas given off. Do the same thing with caustic potash. Mix small quantities of quicklime and ammonium chloride in a mortar, and notice the odor. Has ammonium chloride this odor? Explain the changes that have taken place.

What is the principal source of ammonia compounds?

77. Arrange an apparatus as shown in Fig. 16. In the flask put an intimate mixture of 50 grams of quicklime and 25 grams of ammonium chloride. Allow the inverted funnel to dip just below the surface of the water in the beaker.

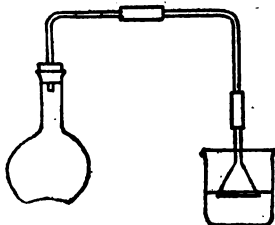


FIG. 16.



FIG. 17.

Heat the flask on a sand bath. After the air is driven out, the gas will be completely absorbed by the water. Disconnect the funnel and turn the delivery tube upward. Bring an inverted gas cylinder over the tube, and collect the escaping gas by displacing the air. As the gas is lighter than air, it is necessary to have the gas jar inverted, and the delivery tube should extend to the bottom. See Fig. 17. The jar

in which the gas is collected should be dry, as water absorbs ammonia very readily. Hence, also, it cannot be collected over water. Into the gas thus collected introduce a burning stick or taper. Does it burn? Does it support combustion? Avoid breathing any quantity of the gas. After enough has been collected, pass the gas into the water as long as it is given off.

78. Try the action of ammonia upon red litmus paper and upon tumeric paper. What effect does it have upon blue litmus paper?

79. Dip a glass rod into strong hydrochloric acid, and bring it in contact with ammonia gas. What is formed? Try the same experiment, using strong nitric acid in place of the hydrochloric acid. What is formed in this case?

Make a tabular statement in your notebook of the physical and chemical properties of ammonia. What is its composition by volume?

Read in a text-book how ammonia is used in making ice.

### NITRIC ACID, $\text{HNO}_3$ .

80. Preparation of Nitric Acid. — In a glass-stoppered retort of 150 to 250 cubic centimeters' capacity put 25 grams of sodium nitrate (Chile saltpeter) and 25 grams<sup>1</sup> of concentrated sulphuric acid. Push over the neck of the retort a glass flask, and allow the latter to dip in cold water contained in the pneumatic trough. See Fig. 8, p. 25. On

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<sup>1</sup> To weigh liquids, balance an empty beaker on the prescription scales; add the required number of weights to the scale pan opposite the one containing the beaker. Then pour the liquid slowly into the beaker until the pan begins to sink.

heating gently, nitric acid will distill over and be condensed in the flask. In the latter stage of the operation the vessel becomes filled with a reddish-brown gas. The acid which is collected has a somewhat yellowish color.

Explain the chemical change that takes place in the retort.

**81. Properties of Nitric Acid.** — Pure nitric acid is a colorless liquid. The colored acid obtained in the preceding experiment can be made colorless by heating it gently in the flask and shaking it. When the acid is boiled, it undergoes slight decomposition into oxygen, water, and compounds of nitrogen and oxygen. One of these is colored, and it is this which is noticed in the last experiment, and whenever strong nitric acid is boiled. Exposed to the light for a long time, a similar decomposition takes place.

In a small flask put a few pieces of granulated tin. Pour on it just enough strong nitric acid to cover it. Heat gently with a small flame. What takes place? What is the appearance of the substance left in the flask? It is a compound of tin, hydrogen, and oxygen. Dissolve a few pieces of copper foil in concentrated nitric acid diluted with about half its volume of water. The operation should be carried on in a flask which has been put in the hood. What action takes place? After it is over, what is the appearance of the liquid in the flask? Pour it out and evaporate to crystallization (be careful not to evaporate to dryness). Compare the substance thus obtained with copper nitrate. Heat specimens of each. Treat small specimens with sulphuric acid. Do the substances appear to be identical? What reasons have you for considering them identical?

## [LABORATORY DEMONSTRATION.]

**82.** Pour fuming nitric acid into a wide test tube so that it is about one fourth full. Heat the end of a stick of charcoal of about the size of a lead pencil and, holding the other end with the forceps, introduce the heated end into the acid. It will continue to burn with a bright light even though placed below the surface of the liquid. The action is oxidation. The charcoal in this case takes oxygen from the acid, and not from the air. Care must be taken in performing this experiment, that the hot charcoal does not touch the sides of the tube; and a large beaker should be placed beneath the test tube, so that in case the tube should break, the acid would be caught and prevented from doing harm.

**83.** Put a small piece of white wool or worsted into an evaporating dish, add a few drops of concentrated nitric acid, and warm gently. Observe the yellow stain. A white feather is stained in the same way. Nitric acid does not stain cotton or calico yellow.

**84.** To a dilute solution of indigo in water add nitric acid. Stir the mixture and keep adding the acid from time to time until the indigo is all bleached. Repeat the experiment, using litmus solution.

**85. Action of Ammonia upon Nitric Acid.** — To a dilute solution of nitric acid in water add ammonia solution until the liquid smells slightly of ammonia, or until a piece of litmus paper dipped into the liquid is colored blue. Now evaporate on the water bath. A solid white crystallized mass remains behind. This is ammonium nitrate, formed by the union of the ammonia with



the nitric acid. Preserve the crystals for future experiments.

**86.** When nitric acid acts upon metals, nitrates are usually formed; for example, nitric acid acting upon copper gives copper nitrate. See Experiment 81. In these compounds the hydrogen of the acid has been replaced by metals. What becomes of the hydrogen? Like the pure acid the nitrates are good oxidizing agents.

Melt a few crystals of potassium nitrate (saltpeter) in an ignition tube, then drop in a bit of charcoal and heat until deflagration takes place.

Repeat the experiment, using a small piece of roll sulphur in place of the charcoal. Explain what takes place.

**87. General Reaction for All Nitrates.** — In a test tube put a small crystal of any nitrate. Add a few drops of pure concentrated sulphuric acid. Cool the mixture, and when cold, pour slowly down the side of the tube a solution of ferrous sulphate. A brown ring is formed where the ferrous sulphate comes in contact with the heavier sulphuric acid solution at the bottom. Ask the instructor for an explanation of the reaction. The experiment is frequently made use of in the analysis of unknown substances.

What are the distinguishing properties of nitric acid? How is it prepared commercially?

### NITROUS OXIDE, $N_2O$ .

**88.** In a glass retort put 10 or 15 grams of ammonium nitrate. Fix the retort in the clamp stand so that the opening of the neck dips under the water of the pneu-

matic trough. Heat the ammonium nitrate gently to secure a regular evolution of the gas. Fill several bottles or jars with the gas as in the experiments under oxygen.

Explain what change takes place in the retort.

89. Insert into a jar of the gas a piece of wood with a live coal on the end. Try also a burning candle. Into another bottle of the gas plunge a bit of burning phosphorus. Also introduce burning sulphur into nitrous oxide. What happens in each case?

90. Fill a cylinder half full of nitrous oxide, and then add hydrogen until full. Apply a light to the mixture: it explodes. What is the reaction? Which of the gases that you have studied does nitrous oxide resemble most nearly? What is the effect of inhaling nitrous oxide?

91. To distinguish nitrous oxide from oxygen, keep a bottle half filled with the gas standing in the pneumatic trough, and use it for Experiment 94, under Nitric Oxide.

### NITRIC OXIDE, NO.

92. Preparation from Copper and Nitric Acid. — Use an apparatus like that used for preparing hydrogen. In the flask put a few pieces of copper foil. Cover this with water. Now add slowly, waiting after each addition, ordinary concentrated nitric acid. When enough acid has been added, gas will be given off. If too much acid is added at once, it not infrequently happens that the evolution of gas takes place so rapidly that the liquid is forced out of the flask through the funnel tube. What is the color of the gas in the flask at first? What is it after the

action has been continued for some time? Collect over water two or three vessels full.

What is the chemical change which causes the formation of nitric oxide?

**93.** Turn one of the vessels containing nitric oxide so that the mouth is upward, and uncover it. What takes place? Explain the appearance of the colored gas in the previous experiment, and the fact that it afterward disappeared. Avoid inhaling the brown fumes that are formed when nitric oxide comes in contact with air.

**94.** Pass nitric oxide from the generating flask into the bottle containing nitrous oxide mentioned under Experiment 91. Are brown fumes formed? If the bottle had contained oxygen, what would have been the result?

**95.** Into one of the bottles containing nitric oxide insert a burning candle. Does the gas burn? Does it support combustion?

**96.** Into a cylinder containing nitric oxide pour a few drops of carbon disulphide,<sup>1</sup> shake the cylinder, and then apply a burning match. A brilliant flame is produced. Substances that evolve a great deal of heat in burning will continue to burn in nitric oxide. When substances burn in nitrous and nitric oxide, what becomes of the nitrogen in these gases?

**97.** Pass a current of nitric oxide, as it issues from the generating flask, into a solution of ferrous sulphate. The

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<sup>1</sup> Be careful to keep the bottle containing carbon disulphide away from the neighborhood of flames. Carbon disulphide vapor mixed with air explodes when lighted.

gas is retained in the solution, and the latter becomes brown, then black in color. If the dark solution is now boiled, the color again becomes yellow or yellowish green. In which previous experiment has this brown solution of nitric oxide in ferrous sulphate been obtained?

In your notebook make a list of the oxides and acids of nitrogen. Give the gravimetric composition of each. What important law do they illustrate? Read in the textbooks how the other oxides are prepared.

### **VOLUMETRIC COMPOSITION OF NITROUS AND NITRIC OXIDES, AND THE LAW OF VOLUMES.**

#### **[LABORATORY DEMONSTRATION.]**

98. Place in a flask of 200 cubic centimeters' capacity 20 grams of potassium nitrate and 80 grams of ferrous sulphate. Add enough water to cover the solids. Close the flask with a tightly fitting rubber stopper carrying a dropping funnel and a delivery tube. Allow concentrated sulphuric acid to enter the flask, drop by drop, and heat gently. A steady stream of pure nitric oxide is evolved. Connect the delivery tube of the flask by means of rubber connections with a T-tube, and connect one of the open ends of the T-tube with a gas-measuring burette full of water, such as was used in the volumetric analysis of the air. Allow the nitric oxide to escape from the open end of the T-tube for 5 or 10 minutes, until all the air has been displaced in the flask and in the connections. Then close the open end of the T-tube with a rubber tube compressed with a pinchcock, and admit the gas into the burette.

Having measured 100 cubic centimeters of the gas in the burette, connect the latter with a simple absorption pipette filled with water, using instead of the ordinary

capillary tube a tube of hard glass  $\frac{1}{8}$  of an inch internal diameter, and filled with granules of metallic copper. (The tube is filled with granular copper oxide, and then the oxide is reduced by heating in a stream of hydrogen.)

The tube containing the copper is heated to a dull red heat by means of a broad-top burner, and the nitric oxide is transferred slowly to the pipette. The gas is thus decomposed into nitrogen, the oxygen uniting with the copper, forming copper oxide. Draw back the gas into the burette, and measure the volume of nitrogen. How does its volume compare with the volume of nitric oxide taken?

Repeat the experiment, using nitrous oxide instead of nitric oxide. Since nitrous oxide is soluble in water, have the water of the burette saturated with this gas before making the measurements. What volume of nitrogen is contained in one volume of nitrous oxide? What is the law of combination of gaseous volumes? Give a number of examples to illustrate the law.

### CHLORINE, Cl.

**99. Preparation of Chlorine.** — A flask of 250 cubic centimeters' capacity provided with a rubber stopper having a funnel tube and a delivery tube is required. The flask used for making hydrogen and nitric oxide may be used. Put into the flask about 100 grams of powdered manganese dioxide, pour upon it enough ordinary concentrated hydrochloric acid to completely cover it. Shake the mixture thoroughly, so that the acid touches all parts of the bottom of the flask. Heat very gently with a small flame, and chlorine will be given off. Collect several dry cylinders or bottles full of chlorine, by letting the delivery tube extend to the bottom of the collecting vessel, and covering

the mouth of the vessel with a piece of paper or a glass plate. You can see when the vessel is full by the color of the gas. [*The experiments with chlorine should be carried on in a good hood. Do not inhale the gas.*]

By what reaction is the chlorine obtained in this experiment?

**100. Properties of Chlorine.** — Notice the color and odor of the gas. Why is it not advisable to collect the gas by the displacement of cold water in the pneumatic trough? Find out from your books how much heavier the gas is than air. Into one of the vessels containing chlorine introduce a little finely powdered antimony. What becomes of the antimony? What causes the evolution of light and heat?

**101.** Into a second jar of chlorine introduce a piece of very thin copper foil or a piece of Dutch leaf. What evidence is there of a transformation of energy? A compound of copper and chlorine is formed.

**102.** Into a third vessel insert a piece of paper with writing on it, some flowers and pieces of colored calico. Most of the colors will be destroyed *if the substances are moist*.

**103.** Into a fourth vessel introduce a dry piece of the same calico as that used in Experiment 102. The dry piece is not bleached, the moist piece is. Explain this.

**104.** Lower a burning candle into a cylinder full of chlorine. See Fig. 18. Why does it burn with a smoky flame? Does the chlorine burn?

**105.** Add chlorine water, made by bubbling chlorine gas through cold water, to a dilute solution of indigo. What effect does it have? Try its action upon litmus solution.

**106.** Moisten a strip of filter paper with turpentine, and plunge it into a bottle of chlorine. See Fig. 19. Turpentine is a compound of carbon and hydrogen. Explain what takes place. Usually so much heat is produced that the substance takes fire. Compare with Experiment 104.



FIG. 18.



FIG. 19.

What are the distinguishing properties of chlorine?

Read in one of the text-books an account of the technical methods of making chlorine, and of the applications of chlorine for bleaching and disinfecting purposes. How can chlorine be liquefied? Describe in your notebooks Faraday's method of liquefying chlorine. Why are oxygen, nitrogen, and hydrogen so much more difficult to liquefy than chlorine, ammonia, and nitrous oxide? What is the *critical* temperature of a gas?

### HYDROCHLORIC ACID, $\text{HCl}$ .

#### [LABORATORY DEMONSTRATION.]

**107.** Light a jet of hydrogen in the air, and carefully introduce it into a vessel containing chlorine. See Fig. 20. Does it continue to burn? What is the appearance of the

flame? What evidence have you that a product is formed? Test the gas remaining in the jar with blue litmus solution shaken up in it, and compare with the action of chlorine on the solution.

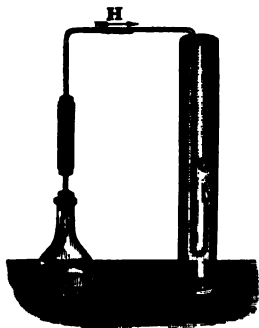


FIG. 20.

What does this experiment show? What is the composition of hydrochloric acid by volume? What law does this illustrate?

#### 108. Preparation of Hydrochloric Acid from Salt and Sulphuric Acid.

—Arrange an apparatus similar to that used in preparing a solution of ammonia in water, but use in place of the one-hole rubber stopper one with two holes, and push a funnel tube through the second hole. Place 50 grams of common salt in the flask. Measure 20 cubic centimeters of water into a beaker, and pour into it 50 cubic centimeters of strong sulphuric acid. When this mixture has cooled down to the ordinary temperature, pour it upon the salt in the flask. Now heat the flask gently, and a regular current of gas will be given off. Conduct it into water; when the air has all been driven out of the flask, the gas is all absorbed by the water under the funnel. The solution of the gas in water is heavier than the water, and as it is formed you can see it settling down in the beaker and mixing with the dilute solution. After the gas has passed into the water for about 10 minutes, disconnect the funnel. Notice the fumes. These become denser by blowing the breath upon them. Apply a lighted match to the end of the tube. Does the gas burn? Collect some of the gas in one or two dry cylinders by downward



displacement of air, as in the case of chlorine. Then connect the generating flask again with the funnel, and let the heating continue until no more gas comes over. There is left in the flask hydrogen sodium sulphate. Write in your notebook the reaction which has taken place.

**109. Properties of Hydrochloric Acid.**—Notice the physical properties of the gas, its color, odor, etc. Insert a burning stick or candle into the gas. Does it support combustion?

**110.** Open another cylinder filled with hydrochloric acid gas under water in the pneumatic trough. If the gas contained no air, the water will rush into the cylinder as it would into a vacuum. What is the reason for this?

**111.** Moisten a strip of filter paper with ammonia solution, and plunge it into a cylinder of hydrochloric acid gas. Explain what takes place.

**112.** Try the action of the solution of hydrochloric acid that you have prepared upon some granulated zinc in a test tube. Is a gas given off? What is it? Add some to a little manganese dioxide, and heat gently. Is chlorine liberated? Add 10 or 12 drops of water to two or three drops of the solution of the hydrochloric acid and taste it. How would you describe the taste? Dip a piece of blue litmus paper into the dilute hydrochloric acid. Litmus is a vegetable color prepared for use as a dye. Other vegetable colors are changed by hydrochloric acid. The color is restored by adding ammonia or caustic soda solution. What is the composition of hydrochloric acid by weight and by volume?

When hydrochloric acid is treated with metals, hydrogen

is evolved, and chlorides of the metals are formed; thus, with zinc, hydrochloric acid gives zinc chloride and hydrogen; with iron, iron chloride and hydrogen, etc. Nearly all the metallic chlorides are soluble in water. Silver chloride, mercurous chloride, and lead chloride are insoluble in water.

**113.** Add a drop of silver nitrate to a solution of any chloride. What is precipitated? What remains in solution?

Add a drop of mercurous nitrate to a solution of any chloride. Interpret the reaction.

Add several drops of lead nitrate to a solution of a chloride. Explain the chemical change.

How is hydrochloric acid obtained technically?

### ACIDS, BASES, AND SALTS.

Acids and bases have the power to destroy the characteristic properties of each other. They *neutralize* each other. The most common acids are sulphuric, nitric, and hydrochloric acids. Among the more common bases are caustic soda, caustic potash, limewater, and ammonia water. Acids turn litmus red, bases turn litmus blue.

**114.** Make a dilute solution of hydrochloric acid by measuring 10 cubic centimeters of dilute hydrochloric acid from the reagent bottle, and diluting it with 50 cubic centimeters of water. Weigh off about 5 grams of sodium hydroxide (caustic soda), and dissolve it in 100 cubic centimeters of water. Measure 20 cubic centimeters of the dilute acid, by means of a burette, into a beaker, add a drop of litmus, and now with another burette graduated into fractions of cubic centimeters run in the caustic

soda solution until the liquid just turns from red to blue (Fig. 21). Note the quantity of caustic soda that was required to neutralize 20 cubic centimeters of the acid. Now take 10 cubic centimeters of acid, and determine the amount of sodium hydroxide solution required to neutralize it. Repeat, using 15 cubic centimeters of acid. What relation do the quantities of alkali used bear to one another? Are they in the proportion of 20 : 10 : 15?

Make a dilute solution of sulphuric acid and of caustic potash as described above, and determine the quantity of alkali required to neutralize 20, 10, and 15 cubic centimeters of the acid.

Does it always require a definite quantity of alkali to neutralize a definite quantity of an acid solution?



FIG. 21.

**115.** When acids neutralize bases, salts and water are formed. Dissolve 10 grams of caustic soda in 100 cubic centimeters of water. Add hydrochloric acid slowly, examining the solution from time to time with a piece of blue litmus paper. As long as the solution is alkaline the paper stays blue; the instant it passes the point of neutralization the paper turns red. When this point is reached, evaporate the solution to complete dryness, and see what is left. Taste the substance; has it an acid taste? If it is common salt or sodium chloride, how ought it to conduct itself when treated with sulphuric acid? Does it conduct itself in this way? Is the substance an alkali or an acid? Is it neutral? Write the equation representing its formation.

**116.** Repeat the experiment, using nitric acid instead of hydrochloric acid. Compare the product with sodium nitrate by treating a small specimen of each with sulphuric acid in test tubes. Are the substances identical? Write the equations representing the neutralization of nitric acid by caustic soda; sulphuric acid by caustic soda; hydrochloric acid by caustic potash; nitric acid by caustic potash. How could you readily determine whether a given substance is an acid, a base, or a salt? Are all salts neutral to litmus?

### **CARBON, C.**

Read the chapter on carbon in your text-book.

In how many forms does the element occur?

What are the characteristic properties of the diamond, graphite, and amorphous carbon? How is charcoal made? How are coke, lampblack, and boneblack or animal charcoal obtained? What is coal? Charcoal can be made in the laboratory by heating wood covered with sand in a crucible until gases are no longer given off.

**117. Decolorizing Power of Boneblack.** — To a dilute solution of indigo add a tablespoonful of animal charcoal. Heat the solution to boiling, shaking frequently; filter while hot. If the filtrate is not colorless, run it through the filter several times. The animal charcoal is very porous, and the coloring matter is retained in the pores.

**118. Chemical Conduct of Carbon.** — At ordinary temperatures carbon is not an active element; it does not unite with hydrogen, oxygen, chlorine, or nitrogen. At high temperatures it readily combines with oxygen. Fasten a bit of charcoal on a wire, heat it in a flame until red hot, then plunge it into a bottle filled with oxygen.

When it stops burning take it out and pour into the bottle a clear solution of limewater. The limewater becomes milky. Explain what has taken place. What is formed when coal burns in a stove or in a grate?

119. At high temperatures carbon will take oxygen from oxides. Mix intimately in a mortar two or three grams of powdered copper oxide and a gram of powdered charcoal. Heat the mixture in an ignition tube provided with a stopper and outlet tube, so that the gas which escapes can be passed into a clear solution of limewater contained in a test tube. Is a white precipitate formed? What evidence have you that oxygen has been extracted from the copper oxide? When cold, pour the black mixture in the tube into cold water in a beaker; stir and let it stand a short time. A red powder will collect on the bottom of the beaker. What is it? Write the equation representing the action of copper oxide upon charcoal at high temperatures. Make a list in your notebook of metals that are obtained from their oxides by the reducing action of carbon.

120. Repeat the preceding experiment, using white arsenic,  $\text{As}_2\text{O}_3$ , in place of the copper oxide. Write the equation representing the reaction.

The element arsenic is volatile, and is hence driven out of the bottom of the tube and deposited on the sides above the mixture in the form of a mirror with a metallic luster.

#### MARSH GAS, $\text{CH}_4$ .

121. Preparation of Marsh Gas.—Put a large spoonful of sodium acetate in an evaporating dish. Heat with a very small flame, and stir until the water of crystallization has been driven off. At first the salt melts, but as the

water is driven off, it again becomes solid. Now weigh off 5 grams of potassium hydroxide,  $7\frac{1}{2}$  grams of quicklime, and 5 grams of dry sodium acetate. Powder the caustic potash and the quicklime in the mortar, then add the acetate, and mix intimately. Put the mixture into an ignition tube, close the tube with a stopper carrying a delivery tube (the apparatus used in making oxygen will do), and support the tube horizontally in a clamp stand. Heat the mixture and collect the marsh gas in the bottles as in the case of oxygen. Notice that the gas is colorless, transparent, and inodorous. It is only slightly soluble in water. Apply a lighted match to some of the gas in a cylinder. Does it burn? Does it give light in burning? What is formed when marsh gas burns?

Read in the text-books about the occurrence and the properties of marsh gas?

### ETHYLENE, $C_2H_4$ .

**122. Preparation of Ethylene.**—Put a few pieces of granulated zinc in a test tube, add alcohol and a drop or two of ethylene bromide. Warm gently. When the gas is coming off rapidly, apply a burning match to the mouth of the tube. Does the ethylene burn with a bright, luminous flame? What is formed when the ethylene burns?

### ACETYLENE, $C_2H_2$ .

**123. Preparation of Acetylene.**—Into a test tube pour 5 cubic centimeters of water, then drop into the water two or three pieces of calcium carbide of about the size of a bean. A rapid evolution of gas takes place. After the air has been driven out of the test tube, apply a light to its mouth. Does the acetylene burn? Is the flame lumi-

nous or non-luminous? Acetylene mixed with air or oxygen explodes violently when a light is applied. Write the equations representing the action of water upon calcium carbide, and of oxygen upon acetylene when it burns in air. Inquire of the instructor how calcium carbide is made.

What weight of carbon is united with one part by weight of hydrogen in marsh gas, ethylene, and acetylene? What law does this illustrate?

### CARBON DIOXIDE, $\text{CO}_2$ .

**124. Carbon Dioxide is formed in Breathing.** — Blow through a glass tube that dips into clear limewater contained in a test tube. What evidence have you that your lungs give off carbon dioxide?

**125. Preparation from Carbonates and Acids.** — To different test tubes containing a little sodium carbonate add dilute hydrochloric, sulphuric, nitric, and acetic acids. Is a gas given off? Test the gas in each tube by bringing into it a wire with a small loop on the end (Fig. 22) that has been dipped into clear limewater. The drop on the end of the wire will get milky if carbon dioxide is present. Treat marble with any acid, and test the gas evolved in the same way. Is it carbon dioxide? Write the reactions in each case.

**126.** Arrange an apparatus like that used in preparing chlorine, or as shown in Fig. 10. Put into the flask some pieces of marble, and pour ordinary hydrochloric acid upon it. Collect the gas by the downward displacement of air, as in the case of chlorine. Fill several bottles with the gas.



FIG. 22.

**127. Properties of Carbon Dioxide.** — Notice whether the gas has color or odor. Insert a lighted candle or burning stick into the gas. What takes place? Is the gas combustible? Does it support combustion? Pour the gas from one of the jars upon the flame of a burning candle; proceed as if pouring water from the jar. See Fig. 23. Pour some of the gas from one vessel to another and show that it has been transferred. Balance a beaker on the prescription scales and pour carbon dioxide into it. If the balance is sensitive, the pan on which the beaker is standing will go down. What do these experiments show?



FIG. 23.

What is the specific gravity of carbon dioxide compared with air? How many times heavier than hydrogen is it? Is it easily liquefied? (Critical temperature?)

**128. Carbon Dioxide acts upon Bases and forms Carbonates.** — Pass a current of carbon dioxide into a solution of caustic soda until it will absorb no more. Add acid to some of the solution thus obtained, and convince yourself that the gas given off is carbon dioxide. Write the equations representing the reactions which take place on passing carbon dioxide into the caustic soda and on adding acid to the solution. Pass carbon dioxide into about 200 cubic centimeters of clear limewater. Filter off the white insoluble precipitate. Try the action of a little acid on it. What evidence have you that it is calcium carbonate? How could you easily distinguish limewater from caustic soda? What are the characteristic properties of carbon dioxide? When carbon is burnt in oxygen, what relation is there between the volume of carbon dioxide and the



volume of oxygen from which it has been formed? Is it larger or smaller?

**129.** Calcium carbonate dissolves in a solution of carbon dioxide in water. Pass carbon dioxide into about 50 cubic centimeters of clear limewater; notice that the precipitate, which at first is very voluminous, gradually goes into solution if you keep on passing in the gas. Acid calcium carbonate,  $\text{H}_2\text{Ca}(\text{CO}_3)_2$ , is formed, and this is soluble. This compound is decomposed on heating the clear solution, and calcium carbonate is again precipitated. Try it.

What causes the "temporary hardness" of natural waters? How can it be removed?

### CARBON MONOXIDE, CO.

**130. Preparation from Oxalic Acid and Sulphuric Acid.**—Put 10 grams of oxalic acid and 50 to 60 grams of concentrated sulphuric acid into a 250 cubic centimeter flask. Close the flask with a stopper provided with a funnel tube and a delivery tube. Connect the delivery tube with two gas washing cylinders containing caustic soda solution. Heat the contents of the flask gently. Carbon dioxide and carbon monoxide are evolved. In passing through the caustic soda the former gas is absorbed. Collect the carbon monoxide in gas bottles over water. Avoid inhaling the carbon monoxide, as it is poisonous. Light some of the gas in a cylinder; notice the characteristic blue flame. What is formed when it burns? By what other methods can carbon monoxide be prepared?

**131. Carbon Monoxide a Reducing Agent.**—Pass carbon monoxide over some heated granules of copper oxide contained in a hard glass tube. Is the oxide reduced? How

do you know? Is carbon dioxide formed? How can it be shown?

In what respects does carbon monoxide resemble and how does it differ from carbon dioxide? How can one be distinguished from the other? In what respects does carbon monoxide resemble and how does it differ from hydrogen?

### THE KINDLING POINT: INFLUENCE OF TEMPERATURE ON CHEMICAL ACTION.

132. Combustible substances must be raised to the kindling temperature before they will burn. When cooled

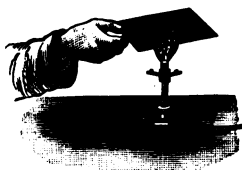


FIG. 24.



FIG. 25.

below this temperature, they are extinguished. Light a Bunsen burner. Bring down upon the flame a piece of wire gauze. There is no flame above the gauze. See Fig. 24. That the gas passes through unburned can be shown by applying a light just above the outlet of the burner and above the gauze. The gas takes fire and burns. By simply passing through the wire gauze the gas is cooled below its burning temperature. Turn on a Bunsen burner. Do not light the gas. Hold a piece of wire gauze about an inch and a half or two inches above the outlet. Apply a lighted match above the gauze, and the gas will burn above the gauze, but not below it. See Fig. 25.

**133.** Take apart a miner's safety lamp and examine its construction. Light it, put it together, and allow a current of coal gas from a rubber tube to strike against the wire gauze. The coal gas will burn inside the lamp, but the flame will not pass through.

What change do the combustibles coal, wood, sulphur, etc., undergo in the air at ordinary temperatures? At elevated temperatures? What influence has temperature upon the rapidity of chemical changes?

### ILLUMINATING GAS.

Read in the text-books how illuminating gas is made. What gases are present in coal gas? How is water gas made, and of what does it consist? Which of the gases present in coal gas and in water gas burn with a luminous, and which with a non-luminous, flame?

### THE BLOWPIPE.

**134.** The flame of the Bunsen burner consists of two parts,—an inner flame of incomplete combustion and an outer flame of complete combustion. In the inner flame there is an excess of combustible substances; these are at a high temperature, and will take away oxygen from certain oxides if they are brought into the flame. The inner flame is called the reducing flame. In the outer flame there is an excess of oxygen, and in this flame certain substances undergo oxidation, hence it is called the oxidizing flame. The blast-lamp and blowpipe flames are similar to the Bunsen burner; each consists of an oxidizing and a reducing flame.

Practice with the blowpipe, and learn to blow without

interrupting the blast on account of breathing. Put a small piece of metallic lead in a depression on a small piece of charcoal and heat it in the oxidizing flame. In making the oxidizing flame, hold the nozzle of the blowpipe in the middle of the gas flame. Notice the formation of lead oxide.

Mix together in a mortar a little tin oxide and sodium carbonate, put some of the mixture into a depression on the charcoal, and heat it in the reducing flame. In making the reducing flame, hold the blowpipe nozzle just outside the gas flame.

On a loop of platinum wire make a borax bead; bring into it a small quantity of any manganese compound. Heat in the oxidizing flame until violet in color. Now heat in the reducing flame until colorless. Repeat this several times.

### THE EQUIVALENT WEIGHTS OF THE ELEMENTS AND THE LAW OF RECIPROCAL PROPORTIONS.

The *equivalent* of an element is that weight of the element which combines with, or replaces in compounds, one part by weight of hydrogen.

**135. Determination of the Equivalent of Zinc.** — Weigh accurately with the analytical balance from .1 to .2 grams of zinc. Place the zinc in a small porcelain crucible, covered with a small inverted funnel. Both crucible and funnel are put into a large beaker nearly full of water, as shown in Fig. 26. A gas measuring tube of 100 cubic centimeters' capacity is filled half full of dilute sulphuric acid, the remaining half of the tube is filled with water, and then the tube is closed with the thumb and inverted

over the crucible and funnel. Hydrogen is evolved and collects in the tube; allow the action to continue until all the zinc has dissolved. [A drop of copper sulphate solution added to the dilute acid hastens the action, but at the same time introduces a slight error.] Transfer the gas-measuring tube to a large cylinder of water, and when it has acquired a constant temperature, read the volume under atmospheric pressure, the temperature, and the height of the barometer. Calculate the weight of hydrogen obtained, and the equivalent of zinc.

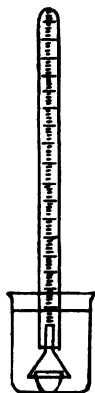


FIG. 26.

**136.** Determine in a similar way the equivalent of magnesium and aluminium. Use hydrochloric acid to dissolve the aluminium. Calculate the equivalent of oxygen from the results obtained in analyzing water. See Determination of the Gravimetric Composition of Water, Experiment 71.

Ask the instructor to explain to you how the equivalents of other elements could be determined. Make a table of equivalent weights in your notebook.

Is the equivalent of an element necessarily a constant quantity? What is the law of reciprocal proportions? Give a number of examples to illustrate it.

### MOLECULAR WEIGHTS.

What is the theoretical explanation given to the laws of definite, multiple, and reciprocal proportions?

What is the hypothesis of Avogadro?

State briefly the principle employed in determining molecular weights. What is the atomic weight of an

element? How is the equivalent of an element related to its atomic weight?

Read in the text-books how the atomic weights are determined.

### **BROMINE, Br.**

**137.** Mix together about a gram of potassium bromide and two grams of manganese dioxide. Pour upon the mixture in a good-sized test tube sufficient dilute sulphuric acid to cover it. Heat gently. What takes place? Interpret the reaction. Compare with the preparation of chlorine.

**138.** Dip a piece of moistened litmus paper into the bromine vapor. Is it bleached? To a dilute solution of indigo add bromine water until the blue color disappears. How does the rapidity of bleaching compare with that of chlorine?

**139.** To a solution of potassium bromide add chlorine water. What is formed? Add a few drops of carbon disulphide, shake, and observe the color of the disulphide. Explain what has taken place.

In what compounds does bromine occur in nature?

Read the chapter on bromine in your text-books, and make a list of the distinguishing properties of the element.

### **HYDROBROMIC ACID, HBr.**

**140. Preparation of Hydrobromic Acid.**—In a test tube put a few crystals of potassium bromide. Pour on them a few drops of concentrated sulphuric acid. Notice the white fumes of hydrobromic acid; at the same time some

reddish brown vapor of bromine is also given off. Treat a few crystals of potassium or sodium chloride in the same way. What difference is there between the two cases? When hydrobromic acid and sulphuric acid decompose one another, what substances are formed?

**141.** The bromides of the metals closely resemble the chlorides; like the latter, they are nearly all soluble in water. Silver, lead, and mercurous bromides are insoluble.

To a solution of potassium bromide add a few drops of a solution of silver nitrate. What is the precipitate? Write the reaction.

To another portion of the potassium bromide add a soluble lead salt. To a third portion add mercurous nitrate solution.

Explain what happens in each case. Compare the results with similar experiments made under Hydrochloric Acid.

How could pure hydrobromic acid gas be prepared? Read descriptions of the method in text-books. Read also about the properties of hydrobromic acid, and notice how it resembles and how it differs from hydrochloric acid. How could one gas be distinguished from the other? What is the volumetric composition of hydrobromic acid?

#### PREPARATION OF AMMONIUM BROMIDE, $\text{NH}_4\text{Br}$ .

**142.** A flask of about 300 cubic centimeters' capacity and a 50 cubic centimeter separating funnel are required (Fig. 27). Measure 110 cubic centimeters of concentrated ammonia solution (sp. gr. .9) into the flask, and, in the hood, measure 37 cubic centimeters of bromine, and pour into the funnel. Support the separating funnel in a clamp stand; under it place the flask with the ammonia. It is well to have

the flask standing in a vessel of cold water. Allow the bromine to run into the ammonia, *drop by drop*. Shake the flask constantly and keep it cool. Nitrogen gas escapes, fumes of ammonium bromide fill the interior of the flask. Do not add bromine after the alkaline reaction has disappeared, or when the liquid becomes yellow add more ammonia. The following equation represents the reaction :

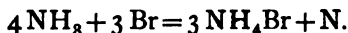


FIG. 27.

Pour the contents of the flask into a large porcelain evaporating dish, and evaporate to dryness on the water bath. The salt can be purified by recrystallization. In your laboratory book, under "Nitrogen," make a note of this method of preparing the gas.

### PREPARATION OF POTASSIUM BROMIDE, KBr.

143. Dissolve 98 grams of ammonium bromide in hot water, and add 100 grams of acid potassium carbonate. Heat to boiling, and continue heating until the odor of ammonia disappears. Then evaporate and crystallize the potassium bromide. The action is thus represented :



Preserve the crystals in a specimen tube.

### IODINE, I.

144. Preparation of Iodine. — Mix about one gram of potassium iodide with about twice its weight of manganese dioxide. Put the mixture in an evaporating dish, add some dilute sulphuric acid, stir to make a pasty mass. Place a clean funnel over the dish and then heat it gently. Grad-



ually the funnel becomes filled with the colored vapor of iodine, and grayish black crystals are deposited on the cold sides of the funnel. Interpret the reaction.

**145.** Melt two or three crystals of iodine in a small test tube and pour the vapor upon a sheet of white paper. Do you obtain crystals or an amorphous powder?

**146. Solubility of Iodine in Solvents.** — Make solutions of iodine in water, in alcohol, and in a water solution of potassium iodide. Use small quantities in test tubes.

**147. Action of Iodine upon Starch.** — Make some starch paste by boiling water in a beaker and adding to it a small quantity of powdered starch mixed with cold water; boil the paste, stirring vigorously with a glass rod. After cooling, add a drop of the water solution of iodine to some of the starch paste. What change takes place? Add starch paste to a dilute water solution of potassium iodide. Is there any change of color? Add now to the last a drop or two of chlorine water. What takes place? Does chlorine alone form a blue compound with starch?

Read about the occurrence, preparation, properties, and uses of iodine in the text-books. How does the specific gravity of iodine vapor compare with that of hydrogen?

### HYDRIODIC ACID, HI.

**148.** Treat a few crystals of potassium iodide with concentrated sulphuric acid in a test tube. What do you notice? Hydriodic acid is more unstable than hydrobromic acid. In the presence of the sulphuric acid it is decomposed; free iodine and the reduction products of sulphuric acid are formed. What reduces the sulphuric acid?

How can pure hydriodic acid be prepared? What are its properties?

**149. Iodides of the Metals Resemble the Chlorides and Bromides.** — Take three test tubes containing dilute solutions of potassium iodide; add silver nitrate to one, lead nitrate to another, mercurous nitrate to the third. Explain what takes place in each case, and compare with the similar experiments under bromides and chlorides.

How can iodides be distinguished from chlorides and bromides?

What is the volumetric composition of hydriodic acid? Compare with hydrochloric and hydrobromic acids?

### **FLUORINE, F, AND HYDROFLUORIC ACID, HF.**

Read about the preparation and properties of fluorine. What is the chief source of fluorine compounds?

**150. Preparation of Hydrofluoric Acid.** — Cover the surface of a piece of glass with paraffine, and with a pointed instrument, such as the end of a file, scratch letters or figures through the paraffine so as to leave the glass exposed where the scratches are made. In a lead dish put 5 or 6 grams of powdered fluorspar, and pour on it enough concentrated sulphuric acid to make a thick paste when stirred up with an iron wire. Put the glass with the paraffined side downward over the vessel containing the fluorspar, and let it stand for several hours. Take off the glass, scrape off the paraffine. The figures which were marked through the coating will be found etched in the glass.

How does the preparation of hydrofluoric acid compare with that of hydrochloric acid? Write the equation. What are the distinguishing properties of hydrofluoric acid?

**SULPHUR, S.**

**151. Properties of Sulphur.** — Heat small pieces of roll sulphur in an ignition tube. Heat very gently at first. Notice that it melts, forming a thin, straw-colored liquid. As the temperature gets higher the liquid becomes darker and darker; after a time it gets so thick that the tube can be turned upside down without danger of its running out. At higher temperatures it becomes more liquid again, and at  $440^{\circ}$  it boils. Having noticed these changes, pour some of the hot sulphur into cold water and examine the plastic sulphur thus obtained.

**152. Monoclinic or Prismatic Sulphur.** — In a covered Hessian crucible or small beaker heat a quantity of roll sulphur. After it has all melted let it cool slowly, and when a thin crust has formed on the surface make a hole through this and pour out the liquid part of the sulphur. The inside of the vessel will be lined with honey-yellow needles. Take out a few crystals and examine them. Are they brittle or elastic? Note their color, and whether they are opaque, transparent, or translucent. Lay the crucible aside, and in the course of a few days again examine the crystals. What changes have taken place?

**153. Rhombic or Octahedral Sulphur.** — Dissolve 2 or 3 grams of roll sulphur in 5 or 10 cubic centimeters of carbon disulphide. Put the solution in a watch crystal, and allow it to evaporate by standing in the air. What is the appearance of the crystals? Are they dark yellow or bright yellow? Are they brittle or elastic? State in tabular form the properties of the allotropic forms of sulphur. Which is the more stable above and below  $96^{\circ}$  C.?

**154.** In an ignition tube heat sulphur to boiling. Introduce into the vapor a strip of thin sheet copper, or hold a narrow strip so that the end just dips into the boiling sulphur. What evidence have you that action takes place? What is formed?

Where is sulphur found in nature? How is it purified and what are its chief uses? Name some of the principal sulphur compounds found in nature.

### HYDROGEN SULPHIDE, $H_2S$ .

**155. Preparation from Ferrous Sulphide,  $FeS$ .**—Into a flask put a small handful of pieces of iron sulphide,  $FeS$ . Close the flask with a stopper provided with a funnel tube and a delivery tube (Fig. 28). Pour dilute sulphuric acid upon the iron sulphide. Pass the gas into water in a test tube. Some of it dissolves, and the water acquires the odor of the gas. Collect a gas jar full of the gas by downward displacement as in the case of chlorine. (Specific gravity 1.178.) Set fire to the gas in the jar. What are the products of combustion? Write the equation representing the formation of hydrogen sulphide.



FIG. 28.

**156.** Pass hydrogen sulphide successively through solutions of lead nitrate, zinc sulphate, and of arsenic chloride (prepared by dissolving a little white arsenic,  $As_2O_3$ , in dilute hydrochloric acid). What do you observe in each case? Sulphides of lead, zinc, and arsenic are formed. Endeavor to write the equation of reaction in each case.

Can sulphur and hydrogen be made to combine directly? How is hydrogen sulphide employed in analytical chemistry? What would be the effect of passing hydrogen sulphide over heated iron? What takes place when steam is passed over heated iron? How can small quantities of hydrogen sulphide be detected?

### SULPHUR DIOXIDE, $\text{SO}_2$ .

**157. Preparation from Copper and Sulphuric Acid.** — Put 8 or 10 pieces of sheet copper, one or two inches long and half an inch wide, into a glass flask. Pour 20 to 30 cubic centimeters of concentrated sulphuric acid upon it. Close the flask with a stopper provided with a delivery tube (Fig. 29). Heat the flask gently and not too rapidly when the gas begins to come off. Saturate water in a bottle with the gas. Collect one or two dry jars full of the gas. This can be readily done by the downward displacement of air, as the gas is more than twice as heavy as air.



FIG. 29.

By what other methods can the gas be obtained? Write the equations. What volume of oxygen is contained in one volume of sulphur dioxide?

**158.** Notice the physical properties of the gas, the appearance, odor, etc. Apply a burning match to the gas in the jar. Does it burn? Does it support combustion? Reserve one bottle of sulphur dioxide for the experiments under sulphuric acid.

**159. Bleaching Action of Sulphur Dioxide.**— Burn sulphur in a gas jar or under a bell jar. Place over the burning sulphur some red flowers (Fig. 30). Allow them to remain in the atmosphere of sulphur dioxide. They will be bleached.



FIG. 30.

**160.** To a solution of rose aniline add some of the solution of sulphur dioxide in water obtained in Experiment 157. What is the effect?

**161.** Neutralize some of the solution of sulphur dioxide with caustic soda. Then evaporate it to dryness. What salt is formed? How is sulphur dioxide prepared, and for what purposes is it used in the arts? Can it be easily liquefied?

**162. General Behavior of Sulphites.**— Sulphites, like carbonates, are decomposed by hydrochloric acid. Pour hydrochloric acid upon solid sodium sulphite in a tube. Notice the effervescence and the odor of the gas.

Nascent hydrogen reduces them, and hydrogen sulphide is evolved. To a solution of sodium sulphite add zinc and hydrochloric acid. The hydrogen sulphide can be recognized by its smell and its blackening action on paper moistened with a solution of a lead salt.

Barium chloride produces a white precipitate of barium sulphite,  $\text{BaSO}_3$ , soluble in hydrochloric acid. This solution, on the addition of chlorine water, yields a white precipitate of barium sulphate, the sulphite being oxidized to sulphate. See reactions for sulphuric acid.

How can sulphur dioxide be made to unite with more oxygen? What are the chief properties of sulphur trioxide?

**SULPHURIC ACID,  $H_2SO_4$ .****163. Preparation from Sulphur, Saltpeter, and Water. —**

Pour a few drops of water into a glass jar and moisten the sides of the vessel, then introduce, by means of the deflagrating spoon, a burning mixture of sulphur and potassium nitrate. Cover the jar, and after letting it stand some time rinse the vessel with distilled water, pouring the water into a test tube. It is a dilute solution of sulphuric acid. Test it with litmus paper and with barium chloride. A white precipitate with the latter reagent consists of barium sulphate,  $BaSO_4$ , and shows that sulphuric acid has been formed.

**164. Preparation from Sulphur Dioxide. —**To a jar containing sulphur dioxide add four or five drops of concentrated nitric acid. Shake the jar, add a little water, and transfer the solution to a test tube and add barium chloride. Interpret the reaction.

**165. Preparation from Sulphur and Nitric Acid. —**Boil a very small quantity of sulphur flowers with concentrated nitric acid in a test tube; then dilute with water and add a few drops of barium chloride solution. What is formed, and what does this experiment show?

How is sulphuric acid manufactured on the large scale? Write out a brief account of the essential features of the process. For what purposes is sulphuric acid used in the arts? What is the specific gravity of concentrated sulphuric acid?

**166. Properties of Sulphuric Acid. It is a Strong Acid.** Add one drop of concentrated sulphuric acid to a large

beaker full of water and test the solution with blue litmus paper.

**167. Heat Evolved when Mixed with Water.**— Into 30 cubic centimeters of water in a beaker pour gradually 120 grams of the concentrated acid, stirring the mixture with a narrow, thin test tube containing a few drops of alcohol, or water. Does the liquid in the tube boil? Determine the temperature of the diluted sulphuric acid with a thermometer.

**168.** Into a test tube containing concentrated sulphuric acid insert a splinter of wood. Why is it blackened? Place about 5 grams of sugar in an evaporating dish and pour 10 cubic centimeters of concentrated acid upon it. Heat gently and stir with a glass rod. Dip a glass rod into dilute sulphuric acid and draw letters on ordinary writing paper, then dry the paper by holding it near the lamp. What do these experiments show?

**169.** Pour 73 cubic centimeters of concentrated sulphuric acid into 27 cubic centimeters of water in a beaker. When cold, transfer the mixture to a measuring cylinder and read the volume. Is it greater or less than 100 cubic centimeters?

What inference can be drawn from this?

**170. Action upon Metallic Oxides.**— Dissolve some powdered copper oxide in dilute sulphuric acid with the aid of heat. Filter the solution, then evaporate it and obtain crystals of copper sulphate. In general, what is formed when an oxide dissolves in sulphuric acid? What is formed when a nitrate is heated with concentrated sulphuric acid? What action does sulphuric acid have upon chlorides? Upon metals?



**171. General Properties of Sulphates.** — To a solution of any sulphate add a solution of barium chloride. What is the change? Is barium sulphate soluble in water and acids?

To a solution of any sulphate add a solution of lead nitrate or acetate. What is precipitated?

Powder a little of the sulphate in a mortar, mix with it sodium carbonate, and heat the mixture in the reducing flame of the blowpipe on charcoal. The sulphate is reduced, and sodium sulphide,  $\text{Na}_2\text{S}$ , is formed. This, when moistened with water on a silver coin, gives a black stain. Treated with dilute acids, the sodium sulphide gives off hydrogen sulphide. This is an exceedingly delicate test for sulphur in any of its compounds.

### PHOSPHORUS, P.

**172.** Examine a stick of phosphorus under water, try to cut it, bend it; observe the color. Take a small piece out of water; notice that it fumes in the air. Dry a small piece between filter paper, and dissolve it in a few cubic centimeters of carbon disulphide in a test tube. Pour the solution upon a large piece of filter paper. Hold the paper with the pincette, and wave it in the air. The disulphide evaporates and leaves a layer of phosphorus finely divided on the paper; this takes fire spontaneously. Take care in making this experiment, and pour any of the solution that may remain into the sink.

**173.** Phosphorus melts easily, and takes fire when heated slightly above its melting point. Float a watch crystal on warm water in a water bath or large beaker. Place a small piece of phosphorus, not larger than a grain of wheat, upon

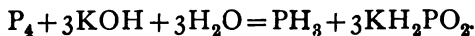
it. Heat the water. The phosphorus melts, and soon after takes fire, burning with a bright flame, forming phosphorus pentoxide,  $P_2O_5$ , and a small quantity of red phosphorus remains behind.

**174.** Bring together in an evaporating dish a very little phosphorus and iodine. What takes place? What is the cause of the heat and the light? What other examples have you had of the direct combination of elements by simple contact? Of direct combination of elements at elevated temperatures?

**175.** Examine red phosphorus; try its solubility in carbon disulphide. Set fire to some of it by heating on an iron plate in the hood. What is formed when it burns? Bring a little red phosphorus in contact with a crystal of iodine. How does the affinity of the red phosphorus compare with that of the yellow variety?

In what forms is phosphorus found in nature? Is it a widely distributed element? Of what importance is calcium phosphate to living things? How is the element obtained in free condition? How is red phosphorus made from yellow phosphorus?

**176. Preparation of Phosphine,  $PH_3$ , or Phosphoretted Hydrogen.**—Drop one or two small pieces of phosphorus into a strong solution of caustic potash contained in a small beaker. Heat the beaker on a sand bath in the hood. After a time bubbles of phosphine are evolved; these take fire spontaneously in the air. Sometimes beautiful smoke rings are formed. Potassium hypophosphite is also formed.



Pure phosphine is not spontaneously inflammable, but some liquid phosphoretted hydrogen,  $P_2H_4$ , is formed at the same time, and this sets fire to the gas. *Phosphine is poisonous.* How many compounds of hydrogen and phosphorus are there? How are phosphorus pentoxide and phosphoric acid prepared? In which experiments have you obtained the former?

**177. General Properties of the Phosphates.**— Make a solution of disodium hydrogen phosphate in water, and to a portion of it add a few drops of silver nitrate solution. A light yellow precipitate of silver phosphate,  $Ag_3PO_4$ , is formed.

To another portion of the phosphate solution add ammonia, ammonium chloride, and magnesium sulphate solution. A white crystalline precipitate of ammonium magnesium phosphate is formed. This reaction is used frequently in testing for phosphoric acid.

To a third portion of the sodium phosphate add nitric acid, and then a solution of ammonium molybdate in nitric acid. A yellow color, and finally a yellow precipitate, is obtained. The reaction is hastened by warming the mixture.

How do monobasic, dibasic, and tribasic acids differ? Give an example of each kind. What acids are obtained from phosphoric acid by heating it to high temperatures? Write the equations.

### ARSENIC, As.

**178.** Examine some metallic arsenic. Notice its luster, and that it is quite brittle. Heat a little of it in an ignition tube. Does it melt? Keep on heating until none is left on the bottom of the tube; what do you notice on the cold parts of the tube? What is this process called?

Heat a little of the element on charcoal before the blow-pipe; it takes fire. What is the color of the flame and of the vapors given off? Notice the peculiar garlic-like odor. What is formed when arsenic burns? Warm a few small crystals with nitric acid. Do they dissolve?

**179.** Mix together about equal small quantities of arsenic trioxide and finely powdered charcoal. Heat the mixture in a small dry tube closed at one end. The arsenic which is set free will be deposited on the walls of the tube in the form of a mirror. Write the equation representing the reaction.

What compounds of arsenic occur in nature?

What is the composition of the substance commonly called "arsenic" or "white arsenic?"

**180.** Heat a small quantity of the trioxide in a dry test tube. What happens? Dissolve some in water to which hydrochloric acid has been added, and conduct hydrogen sulphide into the solution. What is formed? Write the equation.

**181. Preparation of Arseniuretted Hydrogen,  $\text{AsH}_3$ , or Arsine.** — Fit up a flask for generating hydrogen from zinc and dilute sulphuric acid. Connect a calcium chloride U-tube with the delivery tube of the flask, and connect with the U-tube a piece of hard glass tubing drawn out to a small diameter at one end, and turned upward. Put zinc in the flask, and pour dilute sulphuric acid upon it. When the air is out of the vessel, light the hydrogen, and now add slowly a little of a solution of arsenic trioxide,  $\text{As}_2\text{O}_3$ , in dilute hydrochloric acid. What change takes place in the color of the flame? Notice the fumes given off. What are they?

**182.** Hold a cold piece of porcelain—for example, an evaporating dish—in the flame of burning hydrogen and arsine. Notice the appearance of the spots formed. Heat the hard glass tube through which the gas is passing at one point. Just in front of the heated point a thin layer of metallic arsenic will be deposited. This is called a “mirror” of arsenic. What causes this deposit? Minute quantities of arsenic can be detected by the experiments described. It makes no difference in what form of combination the arsenic is put into the flask. This experiment is known as Marsh’s test for arsenic.

**183.** Arsenic acid and the arsenates resemble phosphoric acid and the phosphates in their behavior towards reagents. See General Properties of the Phosphates, Experiment 177. But all arsenic compounds are precipitated by hydrogen sulphide in acidulated solution. Try this with a solution of sodium arsenate acidified with hydrochloric acid.

### ANTIMONY, Sb.

In what form does antimony occur in nature?  
How is the element obtained in free condition?

**184.** Examine a piece of metallic antimony. Notice its white color and brilliant luster. It does not tarnish in the air like arsenic. Heat a small piece on charcoal before the blowpipe. Does it burn? Are fumes formed? Is there an odor as in the case of arsenic? Treat a few small pieces with concentrated nitric acid in a test tube and warm gently; a white powder is formed. ( $\text{Sb}_2\text{O}_3$  and  $\text{HSbO}_3$ .)

Antimony is insoluble in hydrochloric acid.

**185.** Dissolve antimony trioxide in concentrated hydrochloric acid. Pour a few drops into water; notice the formation of a white precipitate,  $\text{SbOCl}$ , "powder of algaroth."

Into another portion of the solution pass hydrogen sulphide; what is formed?

**186. Preparation of Stibine, or Antimoniuretted Hydrogen,  $\text{SbH}_3$ .**—The same apparatus and the same method as that described under Arsine is used for preparing stibine. When hydrogen is being given off, add a solution of tartar emetic or any other antimony compound. Notice the change in the appearance of the flame of hydrogen. Introduce a piece of cold porcelain and notice the antimony deposit. It is darker and more smoky than the arsenic mirror. Heat the hard glass tube and notice the appearance of the antimony mirror in the tube. Is the antimony more or less volatile than the arsenic? For methods of distinguishing the arsenic from the antimony mirror, consult one of the text-books on analytical chemistry.

### **BORON, B.**

**187. Preparation of Boric Acid,  $\text{H}_3\text{BO}_3$ .**—Dissolve 50 grams of powdered borax in 150 cubic centimeters of water. Filter the solution if it is not clear, and then add concentrated hydrochloric acid until the reaction is decidedly acid. Allow the mixture to stand over night. Crystals of boric acid separate. Collect the crystals in a funnel, wash them with a little water, allow them to drain. Dry them by pressing between filter paper. Dissolve some of the crystals in alcohol in a porcelain dish, and set fire to the alcohol. Note the color of the flame, because this reaction is characteristic of boric acid. Write

the equation representing the formation of boric acid from borax. Dip turmeric paper into an aqueous solution of boric acid, then allow it to dry. What is its color?

What compounds of boron occur in nature?

How is the element obtained in the free state?

**188.** Make a bead on platinum wire by melting crystals of boric acid in the loop. Heat with the blowpipe until bubbles of steam are no longer given off. What is the compound that is thus obtained?

What is the composition of borax? How is it used in distinguishing certain metal oxides from one another?

### SILICON, Si.

**189.** To a solution of sodium silicate in water add strong hydrochloric acid. Gelatinous silicic acid is precipitated. Evaporate the mixture of silicic acid and sodium chloride to dryness in an evaporating dish. Moisten the dry mass with hydrochloric acid and again evaporate. Silicon dioxide,  $\text{SiO}_2$ , and sodium chloride are left. Treat with water to dissolve the latter, and filter off and dry the silicon dioxide. Explain the changes that have taken place.

Mention some of the principal varieties of silicon dioxide that occur in nature. In what other forms of combination is silicon found?

How can the element be isolated?

What element does silicon resemble in many respects?

**190. Preparation of Silicon Hydride,  $\text{SiH}_4$ .** — Mix together intimately a saltspoonful of very fine silica and the same quantity of powdered magnesium. Heat the mixture in a dry test tube. Sometimes the action is very violent. Magnesium silicide is formed. When cold, break

the tube and drop fragments of the mass into dilute hydrochloric acid in a beaker. Bubbles of  $\text{SiH}_4$  and hydrogen are given off, which take fire when they come in contact with the air.

**191. Preparation of Silicon Tetrafluoride,  $\text{SiF}_4$ , and Hydrofluosilicic Acid,  $\text{H}_2\text{SiF}_6$ .** — Mix intimately 50 grams of calcium fluoride and 100 grams of sand in a mortar, and pour the mixture into a dry flask. Then add 175 cubic centimeters of concentrated sulphuric acid, and shake the flask so that the acid comes in contact with all parts of

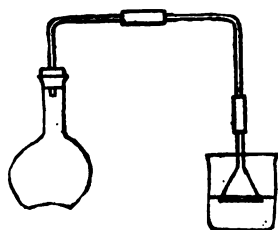


FIG. 31.

the mixture, and with the bottom and sides of the flask. Arrange the apparatus as in Figure 31. The delivery tube and rubber connections must be dry inside. Heat the flask gently on a sand bath. Silicon tetrafluoride is evolved. Pour 200 cubic centimeters of water into the beaker, and allow the funnel to dip but a very little below the surface of the water, and stir the liquid from time to time with a glass rod. When gas is no longer evolved, filter the solution in the beaker. The jelly-like mass is silicic acid. Heat some of it in an evaporating dish and obtain pure silica,  $\text{SiO}_2$ . The clear filtrate contains hydrofluosilicic acid,  $\text{H}_2\text{SiF}_6$ . Test small portions with solutions of potassium chloride and barium chloride. The former gives a gelatinous precipitate of  $\text{K}_2\text{SiF}_6$ ; the latter, a crystalline precipitate of  $\text{BaSiF}_6$ .

Write equations representing the formation of silicon tetrafluoride and hydrofluosilicic acid. How could you show that glass contains silicon?



**CLASSIFICATION OF THE ELEMENTS AND THE PERIODIC LAW.**

In your notebook make a list of the metallic and non-metallic elements you have worked with. In general, how do metals differ from non-metals? Is it possible to satisfactorily classify all the elements under these two divisions? What is meant by a natural family or group of elements? Give examples of such families. Is there any relation between the atomic weights of the elements in these natural families?

What is the system of classification known as the *natural* or *periodic system*? What is the *periodic law*? Read in the text-books about the periodic recurrence of properties in elements, when they are arranged in the order of increasing atomic weight. Make a list of properties that vary periodically. Of what use has the periodic law been in the discovery of new elements? Mention several elements whose existence was foretold by the law.

**POTASSIUM, K.**

Read in the text-books about the occurrence of potassium compounds in nature. Of what importance are potassium compounds to plants? Where does the potassium which plants use come from? When plants are burned, what becomes of the potassium compounds?

**192.** Heat 5 grams of acid potassium tartrate in a small iron evaporating dish, in the hood, until the residue is white. [Any other potassium salt of an organic acid will answer.] Extract the residue with water, filter the solution, and examine it by means of red litmus paper. Is it alkaline?

Examine a solution of potassium carbonate in the same way. Is it alkaline? Evaporate the solution of the residue obtained from the tartrate to dryness, and treat what remains with hydrochloric acid in a test tube. Or you can treat some of the residue obtained by heating the tartrate in a test tube with hydrochloric acid. Test the gas in the tube with a drop of limewater on the loop of wire. What does this experiment show?

How is the metal potassium obtained and what are its chief properties?

**193.** Take a lump of potassium from under the oil; put it on filter paper. Cut off the outside crust; notice the metallic luster and the rapidity with which the metal tarnishes in the air. Throw a small piece not larger than a grain of wheat upon water. What takes place? What is the color of the flame? Is the solution after the action alkaline? Why?

What are the more important compounds of potassium?

**194. Preparation of Potassium Hydroxide, KOH.** — Dissolve about 50 grams of potassium carbonate in 400 to 600 cubic centimeters of water. Heat to boiling in a large iron vessel, and gradually add milk of lime made by adding water to 25 or 30 grams of quicklime. During the operation the mass should be stirred, and water should be added from time to time to prevent the solution from getting too concentrated. Let the liquid cool, and when the precipitate has settled, decant the clear liquid into a bottle. This is a solution of caustic potash. The reaction depends upon the fact that calcium carbonate is insoluble, and potassium carbonate and hydroxide are soluble. Write the equation.

**195.** Examine some solid potassium hydroxide. Allow a small piece to lie exposed to the air for several days. What action does the air have upon it? What is formed when its solution is treated with hydrochloric, nitric, and sulphuric acids? How many salts can it form with sulphuric acid? What are their formulas?

Under what conditions is saltpeter formed? Describe the saltpeter plantations. How is saltpeter used in making nitric acid? In making sulphuric acid?

**196.** Mix together 20 grams of saltpeter, 40 grams of charcoal powder, and 3 grams of sulphur flowers. Put the mixture on an iron plate in the hood, and push a lighted burner under the plate. Stand at some distance. Give a brief account of the manufacture of gunpowder, and explain its action as an explosive.

Nearly all the potassium compounds are soluble in water. The acid potassium tartrate,  $\text{HKC}_4\text{H}_4\text{O}_6$ , and the double chloride of potassium and platinum, and potassium perchlorate are less soluble than the other compounds. Advantage is taken of this fact in testing for the presence of potassium.

**197.** To a moderately strong solution of potassium chloride add a few drops of perchloric acid. The white precipitate is potassium perchlorate,  $\text{KClO}_4$ .

To 50 cubic centimeters of the potassium chloride solution add a solution of tartaric acid, or of sodium bitartrate; allow it to stand after stirring with a glass rod. A white crystalline precipitate of acid potassium

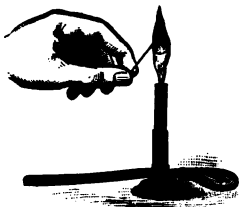


FIG. 32.

tartrate is formed,  $\text{HKC}_4\text{H}_4\text{O}_6$ . Wash a little of the precipitate with alcohol, and test its solubility in hot water. What is its reaction towards litmus paper? Heat some to a high temperature in an evaporating dish. What is formed?

To a drop of potassium chloride solution add several drops of platinum chloride solution. A yellow precipitate,  $\text{K}_2\text{PtCl}_6$ , is formed, soluble in hot water, but insoluble in alcohol and ether.

Dip a platinum wire into the solution of a potassium salt, and bring it into the Bunsen flame. See Fig. 32. Notice the color imparted to the flame. Look at the flame through blue glass. This is a delicate test for potassium compounds.

#### PREPARATION OF POTASSIUM IODIDE, KI.

198. To 3 grams of fine, clean iron filings suspended in 25 cubic centimeters of water in a well-cooled flask, add slowly  $12\frac{1}{2}$  grams of iodine. A slight excess of iron must be used. Filter after standing some time, wash the filter, and to the filtrate add  $2\frac{1}{2}$  grams of iodine. Dissolve 8.2 grams of dry potassium carbonate in boiling water, and to this boiling solution, contained in an evaporating dish, add the solution of iodine and iron. Filter off the precipitate that forms. If the filtrate is not colorless, more potassium carbonate must be added to it. Evaporate the colorless solution of potassium iodide and obtain the crystallized salt.



Pure potassium iodide is colorless, and when its solution is acidified with sulphuric acid, it must not turn starch paste blue. Preserve the crystals in a specimen tube.

**PREPARATION OF POTASSIUM CHLORATE,  $\text{KClO}_3$ .**

**199.** Dissolve 50 grams of potassium carbonate in the smallest possible quantity of hot water. Conduct chlorine gas into the boiling solution until it is no longer alkaline. The chlorine gas must be washed by passing through water before it is conducted into the potassium carbonate solution, otherwise manganese chloride will be carried over, and pink permanganate will be formed. An apparatus arranged on the plan of that shown in Fig. 33 may be used. Dilute the solution of the chlorate with hot water to a volume of 100 cubic centimeters; filter, and allow the salt to crystallize. Use turmeric paper in testing the solution.

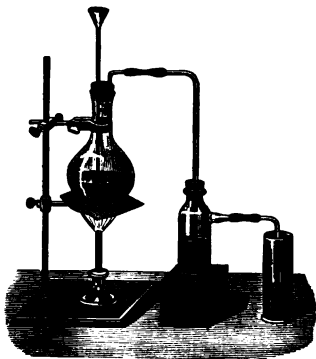


FIG. 33.



Preserve the crystals.

**SODIUM, Na.**

What compounds of sodium are found in nature? How is the free element obtained? Write the equation.

**200.** Examine a piece of metallic sodium; cut off the outside crust; notice the metallic luster. Throw a small piece upon water. What takes place? Throw a small piece upon filter paper floating on the water. Why does the hydrogen take fire in this case? After the action, test the water with red litmus paper. Explain why it is alkaline. Compare with potassium.

What are the chief compounds of sodium?

How is sodium hydroxide obtained?

Describe briefly the processes of making sodium carbonate from salt. For what purposes is soda used?

**201.** The carbonates of sodium and potassium are soluble in water, the carbonates of all other metals, excepting those of rubidium and caesium and some acid carbonates, are insoluble; hence sodium carbonate precipitates solutions of almost all other metals.

To a solution of calcium chloride add a few drops of a solution of sodium carbonate. What is the white precipitate that forms?

Treat a solution of zinc chloride with sodium carbonate. What is formed in this case?

To a solution of ferric chloride add sodium carbonate.

Write the equation representing the reaction in each case.

**202.** Sodium salts are almost all soluble in water. The most characteristic reaction is the bright yellow color which all sodium compounds impart to the Bunsen flame. Dip a clean piece of platinum wire into a solution of a sodium salt, and then bring it into the non-luminous gas flame. Notice the intense yellow color. It is not seen, however, when viewed through a blue glass. It is thus possible to distinguish potassium salts when mixed with sodium salts.

**203. Preparation of Pure Sodium Chloride.** — Make a saturated solution of common salt, 50 grams in 150 cubic centimeters of water, filter the solution, and pass into it hydrochloric acid gas. Allow the gas to pass into the solution through a small funnel which just touches the surface of the liquid. See Fig. 16. Sodium chloride

separates out. Magnesium chloride and other saline impurities remain in solution. Decant the acid solution, wash the residue by decantation with 50 cubic centimeters of cold water, allow it to drain, then heat to dryness with constant stirring in a porcelain dish. Preserve the pure salt in a specimen tube.

### THE SPECTROSCOPE.

**204.** Examine the construction of the spectroscope. Upon what principle is it based?

Observe the continuous spectrum of the luminous flame of the Bunsen burner.

Observe the spectra given by sodium chloride, potassium chloride, and lithium chloride, by dipping clean platinum wires into solutions of these salts, and then bringing them into the non-luminous flame. Compare the colored bands obtained with the colored plates in one of the text-books on chemistry.

Notice also the spectra given by barium, strontium, and calcium chloride solutions.

How can the spectroscope be used in examining substances of unknown composition?

### AMMONIUM SALTS.

Ammonia unites directly with acids to form salts that in many ways resemble the salts of potassium and sodium.

**205.** Place near each other two vessels, one containing strong hydrochloric acid, the other strong ammonia. Blow across the top of these vessels. Explain what you see.

Neutralize dilute hydrochloric acid with a dilute solution of ammonia, and evaporate the solution to dryness on the

water bath. Compare the white salt obtained with potassium and sodium chlorides.

**206.** Heat in a dry test tube one or two crystals of ammonium chloride. Does it sublime, and what is deposited in the cold parts of the tube? The ammonium salts of all volatile acids are volatile, sometimes they can be sublimed, as in the case of the ammonium chloride, otherwise they are decomposed, as in the case of the ammonium salts of the oxygen acids. The salts of non-volatile acids lose their ammonia when fused. Try this with ammonium phosphate. What is formed when ammonium nitrate and ammonium nitrite are heated?

**207.** To a solution of any ammonium salt add caustic soda and then heat the mixture. What is given off? The minutest trace of an ammonium compound can be detected by mixing the substance with dry calcium hydroxide in a small beaker, and covering the beaker with a watch glass, on the under side of which a strip of moist red litmus paper, or moist turmeric paper has been stuck. Try this with a small quantity of sal ammoniac.

**208.** The acid ammonium tartrate and the double chloride of ammonium and platinum, like the corresponding potassium compounds, are soluble with difficulty in water.

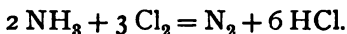
To a solution of an ammonium salt add acid sodium tartrate. What is the formula of the precipitate?

To a concentrated solution of ammonium chloride add a few drops of a solution of platinum chloride. The precipitate is insoluble in alcohol and ether. What is the reaction? How could the ammonium platinum chloride be distinguished from the potassium platinum chloride?



**VOLUMETRIC COMPOSITION OF AMMONIA.**

**209.** Fill a graduated tube, closed at one end and provided with a stopcock of large bore at the other end, with water. Invert the tube over water in the pneumatic trough, and displace the water in the tube by passing in a rapid current of chlorine free from air. Close the stopcock, remove the tube from the water, and pour into the open end 10 cubic centimeters of concentrated ammonia. Open the stopcock cautiously, and allow most of the ammonia to gradually flow down into the part containing the chlorine. Now run in dilute sulphuric acid, in the same way in which the ammonia was run in, to more than neutralize the excess of ammonia. Next fill the tube above the stopcock with water, close the opening with the thumb, and invert over water in a tall cylinder. Open the stopcock, and lower the tube until the gas is under atmospheric pressure. Measure the volume of nitrogen. Three volumes of chlorine should liberate exactly one volume of nitrogen.



Instead of the tube with a glass stopcock of large bore, an ordinary glass tube closed at one end and divided into three parts, by means of rubber bands on the outside, may be used. This is filled with chlorine, and while it is still in the pneumatic trough, it is closed with a tightly fitting rubber stopper carrying a small dropping funnel. Then proceed as before.

**CALCIUM, Ca.**

What are the chief compounds of calcium found in nature? Read in the text-books about the method used in obtaining metallic calcium and the properties of the metal.

**210. Preparation of Calcium Chloride from Calcium Carbonate.** — Dissolve 10 to 20 grams of limestone or marble in ordinary hydrochloric acid. Evaporate the solution to dryness. Heat over the free flame to drive off the water of crystallization. Expose a few pieces of the residue to the air. Does it become moist? In which experiments has calcium chloride been used and for what purpose?

**211. Preparation of Calcium Oxide, or Quicklime.** — Heat a small piece of marble on charcoal to a high temperature with the blowpipe flame. When cold, bring it upon turmeric or red litmus paper, and moisten with a drop of water. Explain what has taken place. What happens when marble is heated to a high temperature in a closed vessel? What is the law of dissociation?

**212.** To 40 or 50 grams of good quicklime add 100 cubic centimeters of water. Afterwards dilute to two or three liters and put the whole in a wide-stoppered bottle. Let it stand several hours, then pour off the clear solution of calcium hydroxide. What takes place when some of the clear solution is exposed to the air? When the breath from the lungs is passed through it? Conduct carbon dioxide into the clear limewater. What is formed? Keep up the current of  $\text{CO}_2$  until the precipitate again dissolves. What is in solution? Boil the clear solution. What is the precipitate that now reappears?

How are stalactites and stalagmites formed?

**213.** Heat some powdered gypsum to about  $200^\circ$  in an air bath. Examine what is left by mixing with a little water so as to form a paste, and let it stand to see whether it will harden. How is plaster of Paris made? Upon what does the hardening depend? What causes the

“permanent hardness” of natural waters? How can it be removed?

**214. General Reactions of Calcium Compounds.**—To a solution of calcium chloride or any calcium compound add sodium carbonate. What is the white precipitate that is formed?

Sulphuric acid precipitates from strong solutions white calcium sulphate, which dissolves in a large excess of water. Try it.

Ammonium oxalate produces, even in dilute solutions, a white precipitate of calcium oxalate, insoluble in acetic acid and in ammonia, but soluble in hydrochloric and nitric acids. Calcium chloride introduced into the non-luminous gas flame gives to it an orange-red color. Try these reactions.

### DETERMINATION OF CARBON DIOXIDE IN CALCIUM CARBONATE.

**215.** Fit up an apparatus as shown in the figure. A small Erlenmeyer flask of about 100 cubic centimeters' capacity is closed with a stopper provided with a calcium chloride tube, and a tube that extends almost to the bottom of the flask. A small specimen tube or test tube is selected, of such size that it can easily be lowered into the flask to take the position shown in Fig. 34. Weigh out accurately .3 to .5 grams of marble in a watch crystal. Transfer the marble without loss to the flask, and pour on it a few cubic centimeters of water so as to close the opening of the long tube. Fill the test tube two thirds full of concentrated hydrochloric acid, and carefully lower



FIG. 34.

it into the flask by means of pincers. Insert the stopper with the calcium chloride tube, and weigh accurately the whole apparatus. Now tip the flask so that the acid flows on the marble. When the action has moderated, warm gently. Finally, when no more gas is given off, displace the carbon dioxide in the apparatus by drawing air through by suction applied at the end of the calcium chloride tube. When the apparatus is perfectly cold, weigh it. The loss in weight of the apparatus is the weight of carbon dioxide in the weight of marble taken. Calculate the per cent.

### **BARIUM, Ba, AND STRONTIUM, Sr.**

What compounds of barium and strontium are found in nature? Barium and strontium compounds closely resemble the compounds of what element?

**216.** The carbonates, sulphates, and oxalates are insoluble in water. Try this by adding to solutions of barium and strontium chlorides, sodium carbonate, sulphuric acid, and ammonium oxalate. In each case white precipitates are formed. Write the equations.

**217.** There are differences in the behavior, however, of these elements, which can be used in detecting their presence when mixed with one another and with other substances. Barium compounds moistened with hydrochloric acid and introduced into the non-luminous flame impart a green color to it. Strontium compounds under similar circumstances give a crimson color.

Hydrofluosilicic acid precipitates solutions of barium salts, but not strontium or calcium solutions.

Strontium sulphate solution precipitates barium, but not strontium or calcium solutions.

Calcium sulphate solution precipitates both barium and strontium solutions, the latter after standing some time, but not calcium solutions.

Potassium chromate precipitates yellow barium chromate insoluble in acetic acid. The same reagent does not precipitate calcium, and strontium is precipitated only in concentrated solutions. The strontium chromate, however, is soluble in acetic acid. Try these reactions.

### MAGNESIUM, Mg.

In what forms does magnesium occur in nature?

Read in the text-books about the preparation of the free element. What happens when the metal is heated in the air? Is it possible to obtain metallic magnesium by the reduction of the oxide with charcoal at high temperatures?

**218.** Magnesium has such a strong affinity for oxygen that it will reduce oxides that are not reduced by carbon. Heat together in a dry test tube as much of a dry mixture of magnesium powder and fine sand as would fill a salt spoon. The action is rather violent. Transfer the product to a small beaker, and add dilute hydrochloric acid. Silicon hydride is evolved, which bursts into flame when it meets the air. Explain what has taken place. What action does magnesium have upon steam at high temperatures?

**219. Characteristic Behavior of Magnesium Compounds.**—To a solution of magnesium chloride add ammonia solution; a white precipitate of magnesium hydroxide,  $\text{Mg}(\text{OH})_2$  is formed. Repeat the experiment, first mixing the magnesium chloride with an equal volume of ammonium chloride solution. No precipitate is formed,

because magnesium hydroxide is soluble in ammonium salts.



An acidulated magnesium solution is not precipitated by ammonia, because ammonium and magnesium salts combine to form double salts, which are not decomposed by ammonia. Try the experiment.

**220.** In the presence of ammonium salts and free ammonia, disodium phosphate,  $\text{HNa}_2\text{PO}_4$ , precipitates magnesium completely as magnesium ammonium phosphate,  $\text{NH}_4\text{MgPO}_4 \cdot 6 \text{H}_2\text{O}$ . The precipitate is at first voluminous, but soon becomes crystalline. In a very dilute solution it forms slowly. The precipitation is hastened by rubbing the glass with a rod. The precipitate is soluble in acids, and is reprecipitated from acid solution by ammonia. This is the reaction most frequently used in testing for magnesium. Try it. Write the equations in your notebook.

**221.** Heat a piece of magnesite intensely before the blowpipe. What change takes place? Moisten with a drop of cobalt nitrate solution and heat again. What change do you notice? (Characteristic Reaction of Magnesium Compounds.)

### ZINC, Zn.

What are the most important zinc ores? How is the metal obtained from its ores? How is it utilized in the laboratory?

**222.** Heat a piece of zinc on charcoal with the blowpipe flame. Notice the crust of zinc oxide. What is its color

while hot, and when cold? Moisten the oxide with cobalt nitrate, and heat again. What color has it now acquired? This is a very characteristic reaction of zinc compounds.

**223.** Alkaline hydroxides precipitate solutions of zinc salts. The zinc hydroxide formed is soluble in excess of the reagent, zincates being formed. Sodium carbonate precipitates basic zinc carbonate, and ammonium sulphide,  $(\text{NH}_4)_2\text{S}$ , gives a white precipitate of zinc sulphide,  $\text{ZnS}$ , soluble in mineral acids, but not in acetic acid. Convince yourself of these facts by experiment, and write the equations representing the reactions.

#### CADMIUM, Cd.

The metal cadmium resembles zinc in appearance. Its compounds accompany the zinc compounds in nature. How is cadmium obtained?

**224.** Heat a small piece of cadmium on charcoal before the blowpipe. A brown incrustation of cadmium oxide is formed. The same coating is obtained by heating a cadmium compound with sodium carbonate on charcoal.

**225.** Pass hydrogen sulphide into a solution of cadmium salt acidified with hydrochloric acid. A yellow precipitate of cadmium sulphide is formed. This is a characteristic reaction for cadmium. Compare with zinc.

#### MERCURY, Hg.

How is mercury obtained? What are its physical properties, such as melting point, boiling point, specific gravity, solubility in water and acids? How many series of salts does it form? How many oxides of mercury are there?

**226.** Mercury is volatilized from all mercury compounds when they are heated with soda. Bring a mixture of dry mercuric chloride and fused soda into a dry glass tube closed at one end, and heat. The volatilized mercury is deposited in the cooler part of the tube as a mirror, or in drops.

**227.** Treat a small drop of mercury with nitric acid that has been diluted with half its volume of water. As long as mercury is in excess, what is formed? Write the reaction. Compare with the action of nitric acid upon copper in the preparation of nitric oxide. How is mercuric nitrate obtained?

**228.** Put in a test tube 5 cubic centimeters of mercurous nitrate, and in another 5 cubic centimeters of mercuric nitrate solution; add slowly hydrochloric acid to each. Notice the difference in the result. Explain what has happened.

In the same way treat mercurous and mercuric solutions with caustic soda. Notice the different behavior. What is formed in each case?

Try also the effect of potassium iodide upon mercurous and mercuric solutions. Filter the red mercuric iodide, and dry it. Then into a clean and dry test tube put a small amount of the mercuric iodide. Warm the middle part of the tube until you cannot bear your hand upon it. Then heat rather gently the lower part of the tube where the mercuric iodide is. Explain what you see. Try the effect of rubbing some of the yellow iodide with a rod. What other substance that you have worked with exists in two crystalline forms, one of which is more stable at ordinary temperatures than the other?



**229.** To solutions of mercurous nitrate and mercuric chloride add hydrogen sulphide. What is formed in each case?

To mercurous nitrate add a solution of stannous chloride. What is the precipitate? Pour off the liquid, and boil the gray precipitate with hydrochloric acid until globules appear.

Try the action of stannous chloride upon mercuric chloride. At first a white precipitate forms, finally a gray one. Explain the reaction. (Characteristic reaction.)

### COPPER, Cu.

In what forms does copper occur in nature? How is it obtained from its ores? Is the metal easily fused? What is formed when it is heated in the air?

What gas is obtained by the action of nitric acid upon it? What gas is obtained by the action of hot sulphuric acid upon it? What is brass? What is bronze?

How many oxides of copper are there? What are their formulas?

**230. Preparation of Cuprous Oxide.**—To a cold dilute solution of copper sulphate add a few cubic centimeters of a solution of grape sugar, then enough caustic potash to dissolve the precipitate which at first forms. Warm the mixture; a yellow precipitate of cuprous hydroxide,  $\text{CuOH}$ , at first forms. Boil, and this becomes red, cuprous oxide,  $\text{Cu}_2\text{O}$ , being formed.

**231. Preparation of Cupric Oxide.**—Place a small quantity of copper nitrate in an evaporating dish or crucible, and heat until fumes are no longer given off. What has taken place?

Put into a test tube a cold solution of copper sulphate, and add to it a solution of caustic soda. What is the precipitate? Repeat the last experiment, but instead of having cold solutions, heat both solutions to boiling, and mix while hot. What is formed in this case?

**232.** Add ammonia, drop by drop, to a solution of copper sulphate, shaking after each addition. Describe what you observe. This is a very delicate test for minute quantities of copper salts.

**233.** Potassium ferrocyanide gives a brown precipitate with copper salts. Add one drop of copper sulphate solution to a test tube full of water, then a drop of acetic acid and a small quantity of potassium ferrocyanide. Describe and explain the change that you observe. This is a very delicate test for copper.

**234.** Saturate 10 cubic centimeters of dilute copper sulphate solution with hydrogen sulphide. Write the reaction.

**235.** Metallic iron and zinc precipitate copper in the metallic state from its solutions. Try this; use a bright wire nail and a bright piece of zinc. What becomes of the zinc and iron?

### **SILVER, Ag.**

In what forms is silver found in nature? How is silver extracted from its ores? Describe the cupellation process of separating silver from other metals. What are the characteristic properties of the metal,—color, malleability, specific gravity, etc.? What are its most important alloys?

**236.** Preparation of Pure Silver Nitrate from a Silver Coin. — Dissolve a 10-cent piece in about 10 cubic centi-

meters of pure concentrated nitric acid diluted with an equal volume of water. Divide the solution into two portions. Evaporate one portion to dryness in an evaporating dish. Heat the dry residue cautiously until it fuses, and the blue copper nitrate is changed to black copper oxide. Do not heat too high, or the silver nitrate also will be decomposed. When cold, treat the residue with water and filter the solution of silver nitrate. Explain the changes that have taken place.

To the second portion add hydrochloric acid as long as a precipitate is formed. Filter, and wash with water until the wash water no longer has an acid reaction. Bring the silver chloride into a porcelain dish, stir it with water, and add caustic soda; heat to boiling and add grape sugar or milk sugar in small pieces. This reduces the silver chloride to metallic silver. Five minutes' boiling and a piece of sugar as large as a hazel nut should suffice. Pour off the liquid, decant a dozen times with water, until the water shows no alkaline reaction. Dissolve the residue in a little warm dilute nitric acid; filter if the solution is not clear, evaporate to dryness on the water bath. Dissolve the crystalline residue in distilled water. Keep the solution of silver nitrate for future use.

**237. General Reactions of Silver Compounds.** — Hydrochloric acid and soluble chlorides give, with silver solutions (use the silver nitrate solution that you have prepared), a white curdy precipitate of silver chloride, insoluble in water and nitric acid, but readily soluble in ammonia, sodium hyposulphite, and potassium cyanide. Caustic soda produces a light brown precipitate of silver oxide,  $\text{Ag}_2\text{O}$ , insoluble in excess of the hydroxide, but soluble in ammonia. Try these reactions.

Hydrogen sulphide or ammonium sulphide produces a black precipitate of silver sulphide,  $\text{Ag}_2\text{S}$ , soluble in potassium cyanide.

Why do silver coins and spoons become black?

Try the action of solutions of sodium carbonate, sodium phosphate, and potassium chromate upon silver nitrate. Explain by means of an equation what happens in each case.

**238.** All silver compounds, when fused with sodium carbonate on charcoal, yield a metallic bead of silver. Try this, using silver chloride.

What action does light have upon silver chloride? How are the silver compounds used in photography?

### ALUMINIUM, Al.

Mention some of the principal compounds of aluminium that occur in nature. How is the metal obtained? What are its most important properties? How does its specific gravity compare with that of the metals in common use?

**239. Action of Acids upon Aluminium.**—Place a small piece of aluminium foil in each of three test tubes, and try the action of hydrochloric, nitric, and sulphuric acids upon it. What happens?

What are the forms of aluminium oxide found in nature?

What is obtained by heating the hydroxide or sulphate of aluminium?

**240.** Heat some aluminium sulphate or ammonium alum upon charcoal before the blowpipe. Notice that the oxide becomes luminous when highly heated, but it does not melt. Moisten it with cobalt nitrate and heat again. What is the color of the mass? (Characteristic reaction.)

Which oxides have you treated with cobalt nitrate in earlier experiments, and what changes were produced?

**241. Preparation and Properties of Aluminium Hydroxide.**—Aluminium hydroxide is a very weak base, and towards strong bases it acts like an acid. To a solution of alum add caustic soda, drop by drop. A gelatinous precipitate appears at first, but finally redissolves when enough caustic soda has been added. (Reactions?) Repeat the experiment, using ammonia instead of caustic soda. Does the precipitate redissolve?

**242.** Sodium carbonate and ammonium sulphide added to a solution of alum precipitate aluminium hydroxide; carbon dioxide is set free in one case, and hydrogen sulphide in the other. What are the equations representing the reactions?

Sodium phosphate precipitates white voluminous aluminium phosphate,  $\text{AlPO}_4$ , when added to a solution of alum. Try these reactions.

**243.** To a small quantity of dilute cochineal solution add an equal bulk of alum solution, and then add ammonia. The flocculent aluminium hydroxide carries down the coloring matter. The colored precipitate is called carmine lake. How is alum used in purifying water?

**244. Preparation of Alum.**—Saturate 10 cubic centimeters of water with powdered aluminium sulphate in a test tube. In another test tube prepare a saturated solution of potassium sulphate. Mix 5 cubic centimeters of each of the solutions. Shake the tube or stir with a rod. Filter, and examine the precipitate with a lens. Is it crystalline?

**TIN, Sn.**

. How does tin occur in nature? How is it prepared commercially?

**245.** Examine a piece of tin foil; notice that it is a white, soft, ductile metal; that it does not change in dry or moist air. Heat a piece on charcoal before the blow-pipe; it melts very easily. At a higher temperature it burns, forming tin dioxide,  $\text{SnO}_2$ . Treat tin with concentrated nitric acid; a white compound, metastannic acid,  $\text{SnO}_3\text{H}_2$ , is formed. Treat tin with concentrated hydrochloric acid, and heat. Stannous chloride is formed.

Mention some of the important alloys of tin. How is tin plate made? Tin forms two oxygen compounds, stannous oxide,  $\text{SnO}$ , with basic properties, and stannic oxide,  $\text{SnO}_2$ , with mainly acid properties.

**246. Behavior of Stannous and Stannic Compounds.** — In each of two test tubes put 1 cubic centimeter of stannous chloride solution; dilute each with 10 cubic centimeters of water. To one of the solutions add enough potassium permanganate solution to give it a faint pink color. Saturate the contents of both tubes with hydrogen sulphide. Notice the different colors of the precipitates. Explain what has taken place. Treat both precipitates with ammonium sulphide. Do they dissolve?

Mercuric chloride gives with stannous solutions at first a white, then a gray, precipitate. Explain. See under Mercury, Experiment 229. Try its action upon stannic compounds. What happens in this case?

**247.** All tin solutions, when treated with metallic zinc, give a precipitate of metallic tin in shining laminæ, or

as a spongy mass. All tin compounds fused with soda and potassium cyanide on charcoal give white globules of metallic tin, and a slight white coating of the dioxide. Try these reactions.

### LEAD, Pb.

What is the chief ore of lead, and how is the metal obtained from it?

**248.** Cut a piece of sheet lead with a pocket knife; notice the color of the bright metal, its softness and ductility. Heat a bit on charcoal. It melts easily; when highly heated, it burns. What is formed? Try its solubility in nitric acid, and in hydrochloric and cold sulphuric acids.

**249. Metallic Lead from Lead Compounds.**—Fuse any lead compound with soda on charcoal. A malleable globule of lead and a yellow coating of lead oxide will be obtained.

Suspend a piece of sheet zinc, as in Fig. 35, in a solution of lead nitrate or acetate, and let it stand overnight. The lead will be deposited in crystallized form. Write the reaction.



FIG. 35.

How many oxides of lead are there, and how are they obtained?

**250.** Treat a little red lead with dilute nitric acid. Filter the solution. What is the brown powder, and what is in solution? Treat some of the brown powder with concentrated hydrochloric acid. What is formed?

**251.** Try the effect of the following reagents upon a solution of a lead salt, such as lead acetate, and explain

what happens in each case : (1) hydrochloric acid or soluble chlorides, (2) sulphuric acid or sulphates, (3) hydrogen sulphide, (4) potassium chromate.

### THE LAW OF SPECIFIC HEATS.

Do different substances have different capacities for heat? What is meant by the *specific heat* of a substance? How can the specific heat of solid substances be determined? Make a table of the specific heats of seven or eight metals, then multiply the specific heat of each metal by its atomic weight. Is the product the same in each case? Suppose that the temperature of 63 lb. of copper, 108 lb. of silver, and 200 lb. of mercury were raised 10 degrees, how would the quantity of heat required to raise the temperature of each metal compare with that required to raise the temperature of the others? What is the law of specific heats? How can it be used to determine atomic weights? Do the specific heats of the elements vary periodically with the atomic weights?

### BISMUTH, Bi.

This element resembles lead in many respects, but it is also related to arsenic and antimony.

**252.** Examine metallic bismuth; notice its reddish luster. It is very brittle. Heat a fragment on charcoal before the blowpipe. What is the color of the oxide that is formed as a coating on the charcoal?

**253.** Dissolve a small quantity of bismuth nitrate in water to which a few drops of nitric acid have been added. Into a portion of the clear solution pass hydrogen sul-



phide ; brownish black bismuth trisulphide is formed. To another portion add potassium chromate ; a yellow precipitate,  $\text{Bi}_2(\text{CrO}_4)_3$ , is formed.

Pour a few drops of the bismuth solution into a large quantity of water ; a white precipitate of basic bismuth salt is formed.

### CHROMIUM, Cr.

What mineral is the chief source of chromium compounds ?

How is potassium chromate obtained from it ?

What are the two important oxides of chromium ?

Which of these is basic and which acid in properties ?

**254. Behavior of the Chromic Salts.** — In these the chromium plays the part of a base. Use a solution of chrome alum for the following experiments. Add caustic soda to the chromium solution, drop by drop. What is formed ? Dissolve the precipitate in an excess of the reagent, then dilute with water and boil. Explain what happens.

Treat portions of the chromium solution with ammonia and ammonium sulphide. Interpret the reactions. How do these reactions of chromium compare with those of aluminium salts ?

**255. Conversion of Chromic Salts into Chromates.** — Fuse in a crucible a mixture of sodium carbonate and saltpeter, and to the fused mass add a small quantity of chromic oxide or chrome alum ; yellow sodium chromate is formed. If a solution of chromium salt is made alkaline with caustic soda, and if an oxidizing agent, such as chlorine water, potassium permanganate, or lead peroxide, is added, the green solution turns yellow, and sodium chromate is formed.

**256. Behavior of Chromates.** — Add acid to the yellow solution of sodium or potassium chromate until it turns red. What is formed? To the red solution add an alkali,  $\text{NaOH}$  or  $\text{Na}_2\text{CO}_3$ , until it is again yellow. Explain the changes that have taken place. To a solution of a chromate add barium chloride. What is the precipitate?

Treat a chromate solution with a soluble lead salt. What is formed?

**257. Preparation of Chromium Trioxide,  $\text{CrO}_3$ .** — Pour 20 cubic centimeters of a saturated solution of potassium dichromate into 30 cubic centimeters of concentrated sulphuric acid. On cooling, red needles of chromic anhydride,  $\text{CrO}_3$ , separate. Allow the crystals to subside, pour off the liquid, then bring some of the crystals upon a porous plate. Try the solubility of some of the crystals in water. Put a few upon filter paper, and observe what effect they have upon it. Pour a drop or two of alcohol upon some of the crystals. The action is violent; sometimes the alcohol takes fire. What is the green substance formed in these experiments?

**258. Conversion of Chromates into Chromic Salts.** — Chromates readily give up oxygen, and are reduced to chromic salts when treated with reducing agents in the presence of an acid. To a solution of potassium dichromate acidified with sulphuric acid, add any one of the following substances, and heat until the color becomes green:  $\text{SO}_2$ ,  $\text{H}_2\text{S}$ ,  $\text{SnCl}_2$ , alcohol, oxalic acid.

What is formed when potassium dichromate is heated with concentrated sulphuric acid? With hydrochloric acid? What gas escapes in each case? What substances are obtained by heating ammonium dichromate?

**MANGANESE, Mn.**

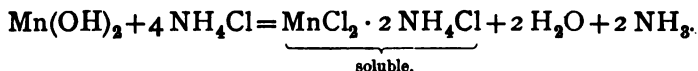
Which compounds of manganese occur in nature? Which of the oxides is commonly used in the preparation of two important non-metallic elements?

The most important manganese compounds besides the dioxide are the manganous salts, the manganates, and permanganates.

**259. Properties of Manganous Salts.** — Make a solution of manganous sulphate, and add to a portion of it ammonium sulphide. What is the color and composition of the precipitate? Try its solubility in acids.

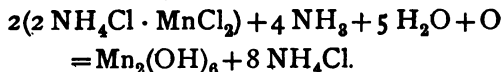
To another portion of the manganous solution add caustic soda. The precipitate rapidly darkens in the air. Explain the changes.

To manganous sulphate add ammonia; a portion of the manganese is precipitated as  $\text{Mn(OH)}_2$ . Manganous hydroxide acts upon ammonium salts, forming soluble double salts. Thus :



Compare this action with that of magnesium hydroxide.

These double salts are not decomposed by ammonia, and hence when ammonia is added to a manganous solution containing free acids or ammonium salts no precipitate is formed. But on standing, oxygen is absorbed from the air, and brown manganic hydroxide is precipitated.



**260. Preparation of Potassium Manganate and Permanganate.** — Any compound of manganese, when fused with alkalis in the presence of oxidizing agents, is converted into an alkaline manganate. Heat 5 grams of KOH with 3 grams of  $\text{KClO}_3$  in a crucible slowly until the mass fuses. Add 5 grams of powdered manganese dioxide, and heat just enough to keep in fusion. Dark green potassium manganate is formed. It dissolves in a *small quantity* of water, with a green color. Pour off the green solution. Add more water to half the solution. Brown manganese perhydroxide,  $\text{MnO}_3\text{H}_2$  or  $\text{Mn}(\text{OH})_4$ , is precipitated, and the color of the solution changes to a beautiful purple, due to the formation of potassium permanganate. Add acid to the rest of the solution. Explain the change.

Owing to the intense color of potassium manganate, the least trace of manganese can be detected by fusing the powdered substance with soda and saltpeter. The smallest quantity of manganese imparts a bluish green color to the fused mass.

**261. Oxidizing Properties of Permanganates.** — To a weak solution of ferrous sulphate, acidified with sulphuric acid, add, drop by drop, a solution of potassium permanganate, until the pink color, which at first disappears, becomes permanent. The ferrous sulphate has been changed to ferric sulphate, and the permanganate to potassium and manganous sulphates. Try to write the equation. Make a dilute solution of oxalic acid. Acidify with sulphuric acid, heat nearly to boiling, and add permanganate as before. Write the equation.

**262. Permanganic Acid.** — The minutest quantity of any manganese compound, when heated with nitric acid and lead peroxide, can be detected by the purple color of the

permanganic acid which is formed. Dilute about 5 cubic centimeters of concentrated nitric acid with an equal volume of water, add about a saltspoonful of red lead or lead peroxide, heat to  $60^{\circ}$  or  $70^{\circ}$ , and now add, *drop by drop*, a weak solution of any manganese salt. The purple color is due to the presence of permanganic acid. This is the most delicate manganese test.

**263.** A small quantity of any manganese compound introduced into a borax bead, and heated in the oxidizing flame, gives an amethyst-colored bead. Heated in the reducing flame, the bead becomes colorless. Try this.

### IRON, Fe.

Name the most important ores of iron. For what purpose is iron pyrites utilized? Describe briefly the method of extracting iron from its ores. What are the differences between cast iron, wrought iron, and steel? How are wrought iron and steel obtained from cast iron?

**264.** Iron is attracted by the magnet; it is infusible before the blowpipe. Finely divided iron, such as powdered iron, takes fire when heated in the air, or when scattered into a flame. An oxide,  $\text{Fe}_3\text{O}_4$ , is formed. With acids it evolves hydrogen, which, owing to contamination with hydrocarbons, has an unpleasant odor.

How many oxides of iron are there? How many classes of iron compounds are there?

**265. Behavior of Ferrous Compounds.**—Treat iron nails or iron filings with hydrochloric acid and warm gently. Ferrous chloride,  $\text{FeCl}_2$ , is formed. Keep iron present in

excess of the acid, and use the solution for the following experiments. To a few cubic centimeters of the solution add caustic soda solution; a white precipitate,  $\text{FeO}_2\text{H}_2$ , is formed, which almost instantly changes, acquiring a dirty green and ultimately a reddish brown color, owing to absorption of oxygen, and conversion into ferric hydroxide,  $\text{Fe}(\text{OH})_3$ .

Add ammonium sulphide to ferrous chloride solution; a black precipitate,  $\text{FeS}$ , ferrous sulphide, is formed. What is the reaction? With potassium ferrocyanide,  $\text{K}_4\text{Fe}(\text{CN})_6$ , a white precipitate,  $\text{K}_2\text{Fe}(\text{CN})_6$ , is formed, which rapidly becomes blue by oxidation to  $\text{Fe}_4(\text{Fe}(\text{CN})_6)_3$ , Prussian blue.

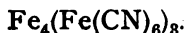
Add potassium ferricyanide solution to ferrous chloride; Turnbull's blue,  $\text{Fe}_8(\text{Fe}(\text{CN})_6)_2$ , is formed. Potassium sulphocyanide with ferrous chloride gives no precipitate if ferric salts are absent. Try it.

**266. Conversion of Ferrous into Ferric Compounds.** — Boil 10 cubic centimeters of ferrous chloride, to which a few drops of concentrated nitric acid have been added, until the color is reddish yellow. Interpret the reaction. Other oxidizing agents, such as potassium permanganate, bromine water, potassium chlorate, and hydrochloric acid, may be used instead of the nitric acid. Reducing agents, such as stannous chloride, zinc and hydrochloric acid, and hydrogen sulphide, convert ferric into ferrous compounds.

**267. Behavior of Ferric Compounds.** — To ferric chloride solution add caustic soda or ammonia. What is formed? Compare with ferrous compounds.

Ammonium sulphide produces black  $\text{FeS}$ , mixed with sulphur.

Potassium ferrocyanide gives Prussian blue,



Potassium ferricyanide with ferric salts in the absence of ferrous salts changes the color of the solution to reddish brown, but does not produce a precipitate.

Potassium sulphocyanide, even in very dilute solutions, produces a blood red color. (Delicate reaction.)

Try all these tests and write the equations representing the changes.

### NICKEL, NI, AND COBALT, Co.

Nickel and cobalt are white metals. They resemble iron. Like iron, they are attracted by a magnet. They dissolve slowly in hydrochloric and sulphuric acids, and readily in nitric acid.

**268.** Make solutions of nickel and cobalt nitrates in water. Add ammonium sulphide to each solution, and explain what takes place. In the same way, try the effects of ammonia, caustic soda, and potassium cyanide solutions upon the nickel and cobalt solutions. Notice the different behavior. Add potassium nitrite to solutions of nickel and cobalt acidified with acetic acid. In the case of cobalt a yellow precipitate is obtained.

## TABLE OF THE ELEMENTS

WITH SYMBOLS AND ATOMIC WEIGHTS.

NAME.	SYMBOL.	ATOMIC WEIGHT.	NAME.	SYMBOL.	ATOMIC WEIGHT.
Aluminium .....	Al	27.	Neodymium .....	Nd	140.5
Antimony .....	Sb	120.	Nickel .....	Ni	58.7
Arsenic .....	As	75.	Nitrogen .....	N	14.03
Barium .....	Ba	137.	Osmium .....	Os	190.8
Bismuth .....	Bi	208.9	Oxygen .....	O	16.
Boron .....	B	11.	Palladium .....	Pd	106.6
Bromine .....	Br	79.95	Phosphorus .....	P	31.
Cadmium .....	Cd	112.	Platinum .....	Pt	195.
Cæsium .....	Cs	132.9	Potassium .....	K	39.11
Calcium .....	Ca	40.	Praseodymium .....	Pr	143.5
Carbon .....	C	12.	Rhodium .....	Rh	103.
Cerium .....	Ce	140.2	Rubidium .....	Rb	85.5
Chlorine .....	Cl	35.45	Ruthenium .....	Ru	101.6
Chromium .....	Cr	52.1	Samarium .....	Sm	150.
Cobalt .....	Co	59.	Scandium .....	Sc	44.
Columbium .....	Cb	94.	Selenium .....	Se	79.
Copper .....	Cu	63.6	Silicon .....	Si	28.4
Erbium .....	Er	166.3	Silver .....	Ag	107.92
Fluorine .....	F	19.	Sodium .....	Na	23.05
Gadolinium .....	Gd	156.1	Strontium .....	Sr	87.6
Gallium .....	Ga	69.	Sulphur .....	S	32.06
Germanium .....	Ge	72.3	Tantalum .....	Ta	182.6
Glucinum .....	Gl	9.	Tellurium .....	Te	125.
Gold .....	Au	197.3	Terbium .....	Tb	160.
Hydrogen .....	H	1.007	Thallium .....	Tl	204.18
Indium .....	In	113.7	Thorium .....	Th	232.6
Iodine .....	I	126.85	Thulium .....	Tu	170.7
Iridium .....	Ir	193.1	Tin .....	Sa	119.
Iron .....	Fe	56.	Titanium .....	Ti	48.
Lanthanum .....	La	138.2	Tungsten .....	W	184.
Lead .....	Pb	206.95	Uranium .....	U	239.6
Lithium .....	Li	7.02	Vanadium .....	V	51.4
Magnesium .....	Mg	24.3	Ytterbium .....	Yb	173.
Manganese .....	Mn	55.	Yttrium .....	Yt	89.1
Mercury .....	Hg	200.	Zinc .....	Zn	65.3
Molybdenum .....	Mo	96.	Zirconium .....	Zr	90.6



## TABLE OF THE WEIGHTS OF GASES

UNDER STANDARD CONDITIONS (0° C. AND 760 MM.).

*Weight of 1 Liter in Grams.*

Air	.	.	.	—	.	.	.	1.29327
Oxygen	.	.	.	O <sub>2</sub>	.	.	.	1.4295
Nitrogen	.	.	.	N <sub>2</sub>	.	.	.	1.2572
Hydrogen	.	.	.	H <sub>2</sub>	.	.	.	.0900

To find the weight of a liter of any other gas, multiply its specific gravity compared with air as the unit by the weight of a liter of air, or multiply half the molecular weight of the gas by the weight of a liter of hydrogen.

$$\text{Wt. of 1 Liter} = \frac{1}{2} \times \text{Mol. Wt.} \times .0900.$$

## TABLE OF TENSION OF AQUEOUS VAPOR

IN MM. OF MERCURY.

t° C.	MM.	t° C.	MM.	t° C.	MM.
-10	2.08	11	9.79	40	54.91
- 9	2.26	12	10.46	45	71.39
- 8	2.46	13	11.16	50	91.98
- 7	2.67	14	11.91	55	117.48
- 6	2.89	15	12.70	60	148.79
- 5	3.13	16	13.54	65	186.94
- 4	3.39	17	14.42	70	233.08
- 3	3.66	18	15.36	75	288.50
- 2	3.96	19	16.35	80	354.62
- 1	4.27	20	17.39	85	433.00
0	4.60	21	18.50	90	525.39
1	4.94	22	19.66	95	633.69
2	5.30	23	20.89	99	733.21
3	5.69	24	22.18	100	760.00
4	6.10	25	23.55	101	787.59
5	6.53	26	24.99	105	906.41
6	7.00	27	26.51	110	1075.37
7	7.49	28	28.10		
8	8.02	29	29.78		
9	8.57	30	31.55		
10	9.17	35	41.83		

# INDEX.

	PAGE		PAGE
Acetylene . . . . .	56	Cadmium . . . . .	97
Acids, Bases, and Salts . . . .	52	Calcium . . . . .	91
Air, Properties of . . . . .	11	Calcium, Reactions of . . . .	93
Air, Volumetric Composition of .	23	Carbon . . . . .	54
Aluminium . . . . .	102	Carbonates, Reactions of . . .	57
Ammonia . . . . .	39	Carbon Dioxide . . . . .	57
Ammonia, Volumetric Composition		Carbon Dioxide, Determination of	93
of . . . . .	91	Carbon Monoxide . . . . .	59
Ammonium Bromide, Preparation		Chemical and Physical Changes .	3
of . . . . .	65	Chlorides, Reactions of . . . .	52
Ammonium Compounds, Reac-		Chlorine . . . . .	47
tions of . . . . .	89	Chromates . . . . .	108
Antimoniuretted Hydrogen . . .	80	Chromium . . . . .	107
Antimony . . . . .	79	Classification of Elements . .	83
Apparatus, Construction of . . .	1	Cobalt . . . . .	113
Aqueous Vapor, Tension of . . .	116	Compounds, Chemical, defined .	9
Arsenates, Reactions of . . . .	79	Conservation of Matter . . . .	20
Arsenic . . . . .	77	Copper . . . . .	99
Arsine, Arseniuretted Hydrogen	78	Critical Temperature . . . . .	49
Atomic Weights, Table of . . .	114		
Balance, The Chemical . . . . .	10	Definite Proportions, Law of . .	21
Barium . . . . .	94	Diffusion, Law of . . . . .	30
Bases, Acids, and Salts . . . .	52	Dissociation, Law of . . . . .	92
Bismuth . . . . .	106		
Blowpipe, The . . . . .	61	Elements and Compounds defined	9
Boric Acid . . . . .	80	Energy, Conservation of . . . .	20
Boron . . . . .	80	Equivalent Weights of Metals .	62
Bromides, Reactions of . . . .	65	Ethylene . . . . .	56
Bromine . . . . .	64	Fluorine . . . . .	68

	PAGE		PAGE
Gas Volumes, Law of . . . .	46	Matter and Energy . . . .	7
Gases, Law of Diffusion of . .	30	Measurement of Gases . . .	14
Gases, Measurement of . . .	14	Measurement of Liquids . . .	10
Gases, Weights of . . . .	115	Mercury . . . . .	97
Homogeneous Matter . . . .	9	Mixtures, Mechanical . . . .	8
Hydriodic Acid . . . . .	67	Molecular Weights . . . . .	63
Hydrobromic Acid . . . . .	64	Multiple Proportions, Law of .	37
Hydrochloric Acid . . . . .	49	Nickel . . . . .	113
Hydrofluoric Acid . . . . .	68	Nitrates, Reaction of . . . .	43
Hydrofluosilicic Acid . . . .	82	Nitric Acid . . . . .	40
Hydrogen . . . . .	26	Nitric Oxide . . . . .	44
Hydrogen Dioxide . . . . .	37	Nitric Oxide, Volumetric Compo-	
Hydrogen Sulphide . . . . .	70	sition of . . . . .	46
Illuminating Gas . . . . .	61	Nitrogen . . . . .	22
Iodides, Reactions of . . . .	68	Nitrous Oxide . . . . .	43
Iodine . . . . .	66	Oxygen . . . . .	16
Iron . . . . .	111	Ozone . . . . .	36
Kindling Point, The . . . .	60	Periodic Law . . . . .	83
Law of Conservation of Matter .	20	Permanganates . . . . .	110
Law of Definite Proportions . .	21	Phosphates, Reactions of . . .	77
Law of Diffusion of Gases . . .	30	Phosphine . . . . .	76
Law of Dissociation . . . . .	92	Phosphorus . . . . .	75
Law of Gas Volumes . . . . .	46	Physical and Chemical Changes .	3
Law of Multiple Proportions . .	37	Potassium . . . . .	83
Law, Periodic . . . . .	83	Potassium, Reactions of . . .	85
Law of Reciprocal Proportions .	62	Potassium Bromide, Preparation	
Law of Specific Heats . . . .	106	of . . . . .	66
Lead . . . . .	105	Potassium Chlorate, Analysis of .	21
Liquids, Measurement of . . .	10	Potassium Chlorate, Preparation	
Liquefaction of Gases . . . .	49	of . . . . .	87
Magnesium . . . . .	95	Potassium Iodide, Preparation of	86
Manganates . . . . .	110	Rapidity of Chemical Action . .	60
Manganese . . . . .	109	Reciprocal Proportions, Law of .	62
Marsh Gas . . . . .	55	Salts, Acids, and Bases . . . .	52
Marsh's Test for Arsenic . . .	78	Silicic Acid . . . . .	81
Matter, Conservation of . . . .	20	Silicon . . . . .	81
Matter, Homogeneous . . . .	9	Silicon Hydride . . . . .	81
		Silicon Tetrafluoride . . . .	82

	PAGE		PAGE
Silver . . . . .	100	Table of Elements . . . . .	114
Silver Nitrate, Preparation of . . . . .	100	Table of Tension of Aqueous	
Sodium . . . . .	87	Vapor . . . . .	116
Sodium Chloride, Preparation of . . . . .	88	Table of Weights of Gases . . . . .	115
Specific Gravity . . . . .	10	Tin . . . . .	104
Specific Heats, Law of . . . . .	106	Volumes, Law of Gas . . . . .	46
Spectroscope . . . . .	89	Water, Properties of . . . . .	24
Stibine . . . . .	80	Water, Composition by Volume	
Strontium . . . . .	94	and by Weight . . . . .	32
Sulphates, Reactions of . . . . .	75	Zinc . . . . .	96
Sulphites, Reactions of . . . . .	72	Zinc, Equivalent of . . . . .	62
Sulphur . . . . .	69		
Sulphur Dioxide . . . . .	71		
Sulphuric Acid . . . . .	73		









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